

LABORATORY KEYS TO CHEMISTRY

University Of Alberta



0 0000 12479 31

ELAINE W. LEDBETTER • JAY A. YOUNG



CURRICULUM

QD
33
L46
1973
LAB.MAN.

CURR

Ex libris
UNIVERSITATIS
ALBERTAENSIS



LABORATORY KEYS TO CHEMISTRY

ELAINE W. LEDBETTER • JAY A. YOUNG

▲ ADDISON-WESLEY PUBLISHING COMPANY

Menlo Park, California • Reading, Massachusetts • London • Don Mills, Ontario

LIBRARY
UNIVERSITY OF ALBERTA

COMPONENTS OF THE KEYS TO CHEMISTRY PROGRAM

Text

Laboratory Manual

Teachers' Guide

Tests

PHOTO AND ILLUSTRATION CREDITS

Illustrations by Liz James

Photos by Jon Emerson unless otherwise indicated

Photos pp. 16 and 56 by Ernest A. Umemoto

Copyright © 1973 by Addison-Wesley Publishing Company, Inc.
Phillippines Copyright 1973

All rights reserved. No part of this publication may be reproduced, stored in a retrieval system, or transmitted, in any form or by any means, electronic, mechanical, photocopying, recording, or otherwise, without the prior written permission of the publisher. Printed in the United States of America. Published simultaneously in Canada.

ABCDEFGHIJK76543

**LIBRARY
UNIVERSITY OF ALBERTA**

TABLE OF CONTENTS

Table of Atomic Weights v

KEYS TO THE CHEMISTRY LABORATORY

The Importance of Laboratory Work to Chemistry 1

Preparing for Laboratory Work 2

Rules of Conduct in the Laboratory 3

Safety in the Chemistry Laboratory 4

LABORATORY TECHNIQUES AND PROCEDURES

The Gas Burner 5

Measuring Weight 7

Measuring Liquid Volume 8

Transferring Solids from Reagent Bottles 10

Transferring Liquids from Reagent Bottles 12

Filtering Precipitates 13

Recovering a Solid from a Liquid 16

Pipetting a Liquid 17

Working with Glass 18

Experiment

- 1-1 Practice in Scientific Observation 21
- 1-2 Some Useful Laboratory Techniques 23
- 1-3 Learning to Work with Variables 27
- 1-4 Measurement and the Concept of Density 31
- 1-5 Heat Energy and Randomness 33

Exercise

- 1-1 Learning to Use a Chemistry Handbook 35

Experiment

- 2-1 Classification of Matter 37
- 2-2 Physical and Chemical Properties of Matter 39
- 2-3 Melting and Freezing of Solids 41
- 2-4 Elements, Mixtures, and Compounds 43
- 2-5 Forces between Like and Unlike Magnetic Poles 45
- 3-1 Learning to Recognize Dynamic Equilibrium 47
- 3-2 The Process of Boiling 49
- 3-3 The Chemical Properties of Water 51
- 3-4 Energy Changes in Solutions 53
- 3-5 Unsaturated, Saturated, and Supersaturated Solutions 55
- 4-1 Electrical Conductivity of Liquids 57
- 4-2 Preparation of a Salt from a Typical Acid and Base 59
- 4-3 Acid-Base Indicators 61
- 4-4 Hydrolysis, the Common Ion, and Buffers 63

Experiment

- 5-1 Building a Model from Indirect Evidence 66
- 5-2 Identification of Ions by Color 67
- 5-3 Paper Chromatography 71

- 6-1 Grouping Elements into Families on the Basis of Similar Electron Patterns 73
- 6-2 The Relationship between Atomic Number and the Radius of Atoms 75
- 6-3 The Relationship between Atomic Number and Ionization Energy 77
- 6-4 A Convenient Method for Grouping Elements 79
- 6-5 Separation of Group I Cations 80
- 6-6 Identification of Group II Cations 82

- 7-1 The Relationship between Bond Types and the Solution Process 84
- 7-2 The Relationship between Bond Types and Physical Properties 86
- 7-3 Energy in Chemical Reactions 88

- 8-1 Predicting Reaction Rates 90
- 8-2 The Effect of Concentration and Temperature on Reaction Rate 91
- 8-3 Types of Chemical Reactions 94
- 8-4 A Quantitative Study of a Chemical Reaction 97

- 9-1 Diffusion of Gases 100
- 9-2 The Effect of Pressure on the Volume of Gases 102
- 9-3 The Effect of Temperature on the Volume of a Gas 104
- 9-4 The Molar Volume of a Gas 106
- 9-5 Testing Predictions in the Laboratory 109
- 9-6 Determination of the Formula for an Unknown Compound 110

- 10-1 Water in Crystals 113
- 10-2 The Structure of Crystals 115
- 10-3 Ionic Reactions in Water Solutions 118

- 11-1 Identification of Certain Anions 122
- 11-2 Elementary Qualitative Analysis 126

APPENDICES

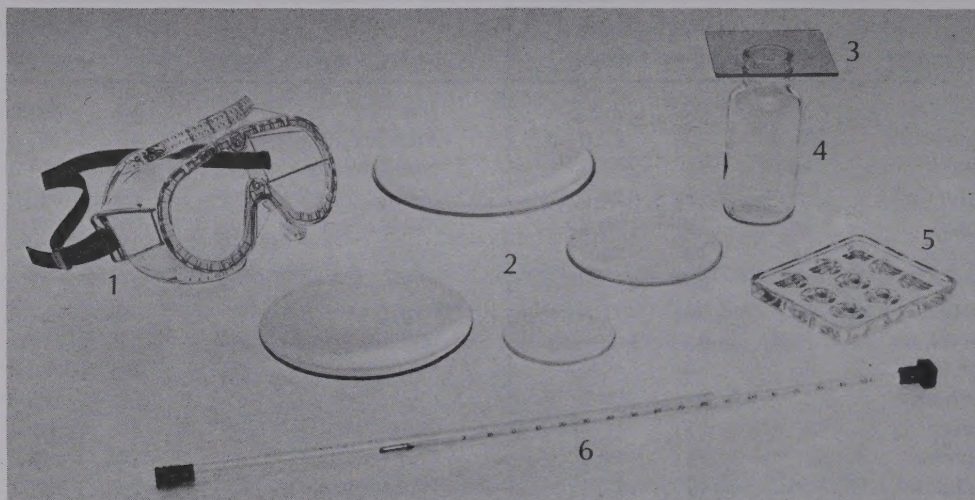
- Vapor Pressure of Water 132
- Charges and Oxidation Numbers of Common Ions 133
- Periodic Table of the Elements 134
- Periodic Table of Electronegativities 136
- Periodic Table of Oxidation Numbers 137

INDEX 138

TABLE OF ATOMIC WEIGHTS

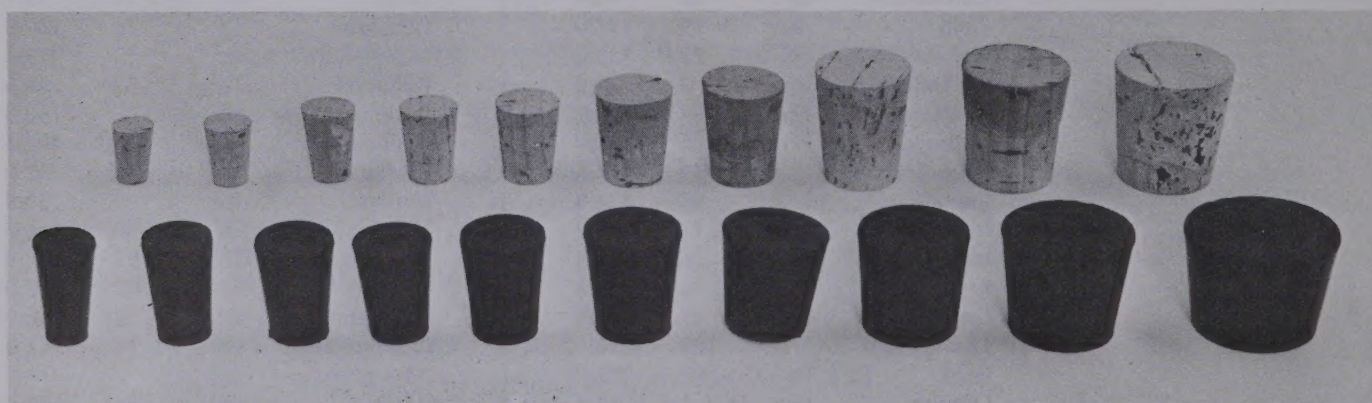
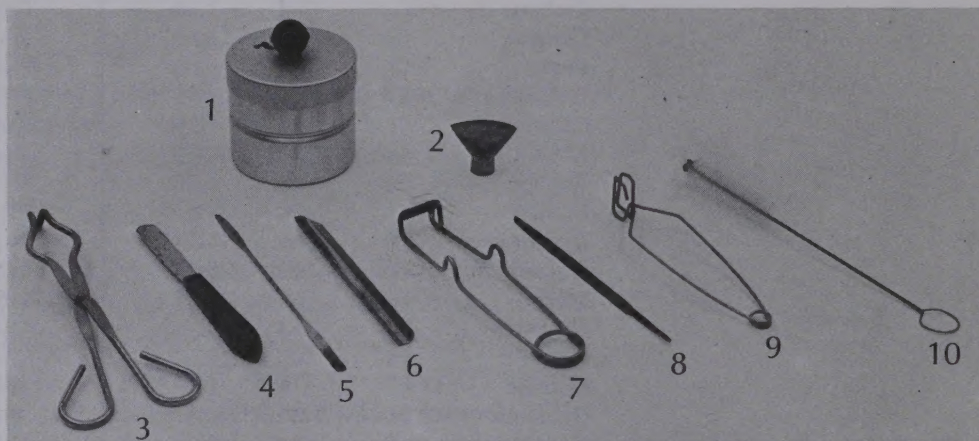
Element	Symbol	Atomic number	Atomic weight	Element	Symbol	Atomic number	Atomic weight
Actinium	Ac	89	(227)	Mercury	Hg	80	200.6
Aluminum	Al	13	27.0	Molybdenum	Mo	42	95.9
Americium	Am	95	(243)	Neodymium	Nd	60	144.2
Antimony	Sb	51	121.8	Neon	Ne	10	20.2
Argon	Ar	18	40.0	Neptunium	Np	93	(237)
Arsenic	As	33	74.9	Nickel	Ni	28	58.7
Astatine	At	85	(210)	Niobium	Nb	41	92.9
Barium	Ba	56	137.3	Nitrogen	N	7	14.0
Berkelium	Bk	97	(249)	Nobelium	No	102	(254)
Beryllium	Be	4	9.0	Osmium	Os	76	190.2
Bismuth	Bi	83	209.0	Oxygen	O	8	16.0
Boron	B	5	10.8	Palladium	Pd	46	106.4
Bromine	Br	35	79.9	Phosphorus	P	15	31.0
Cadmium	Cd	48	112.4	Platinum	Pt	78	195.1
Calcium	Ca	20	40.1	Plutonium	Pu	94	(242)
Californium	Cf	98	(251)	Polonium	Po	84	(210)
Carbon	C	6	12.0	Potassium	K	19	39.1
Cerium	Ce	58	140.1	Praseodymium	Pr	59	140.9
Cesium	Cs	55	132.9	Promethium	Pm	61	(145)
Chlorine	Cl	17	35.5	Protactinium	Pa	91	(231)
Chromium	Cr	24	52.0	Radium	Ra	88	(226)
Cobalt	Co	27	58.9	Radon	Rn	86	(222)
Copper	Cu	29	63.5	Rhenium	Re	75	186.2
Curium	Cm	96	(247)	Rhodium	Rh	45	102.9
Dysprosium	Dy	66	162.5	Rubidium	Rb	37	85.5
Einsteinium	Es	99	(254)	Ruthenium	Ru	44	101.1
Erbium	Er	68	167.3	Samarium	Sm	62	150.4
Europium	Eu	63	152.0	Scandium	Sc	21	45.0
Fermium	Fm	100	(252)	Selenium	Se	34	79.0
Fluorine	F	9	19.0	Silicon	Si	14	28.1
Francium	Fr	87	(223)	Silver	Ag	47	107.9
Gadolinium	Gd	64	157.3	Sodium	Na	11	23.0
Gallium	Ga	31	69.7	Strontium	Sr	38	87.6
Germanium	Ge	32	72.6	Sulfur	S	16	32.1
Gold	Au	79	197.0	Tantalum	Ta	73	180.9
Hafnium	Hf	72	178.5	Technetium	Tc	43	(99)
Helium	He	2	4.0	Tellurium	Te	52	127.6
Holmium	Ho	67	164.9	Terbium	Tb	65	158.9
Hydrogen	H	1	1.0	Thallium	Tl	81	204.4
Indium	In	49	114.8	Thorium	Th	90	232.0
Iodine	I	53	126.9	Thulium	Tm	69	168.9
Iridium	Ir	77	192.2	Tin	Sn	50	118.7
Iron	Fe	26	55.8	Titanium	Ti	22	47.9
Krypton	Kr	36	83.8	Uranium	U	92	238.0
Lanthanum	La	57	138.9	Vanadium	V	23	50.9
Lawrencium	Lr	103	(256)	Wolfram	W	74	183.9
Lead	Pb	82	207.2	(Tungsten)			
Lithium	Li	3	6.9	Xenon	Xe	54	131.3
Lutetium	Lu	71	175.0	Ytterbium	Yb	70	173.0
Magnesium	Mg	12	24.3	Yttrium	Y	39	88.9
Manganese	Mn	25	54.9	Zinc	Zn	30	65.4
Mendelevium	Md	101	(256)	Zirconium	Zr	40	91.2

Note: The atomic weight for each element is based on the carbon-12 scale. Parentheses denote atomic weight for most stable or best known isotope.



(1) Eye protectors (these should be worn any time that there is question of danger to your eyes), (2) watch glasses of several sizes, (3) cover glass, (4) gas-collecting bottle, (5) spot plate, (6) thermometer with transparent storage case.

An assortment of small equipment items commonly used in chemistry laboratories. (1) Desiccator, (2) wingtop flame spread for the gas burner, (3) crucible tongs, (4), (5), (6) three kinds of spatulas, (7) striker for lighting burners, (8) triangular file, (9) test tube holder, (10) brush for cleaning test tubes.

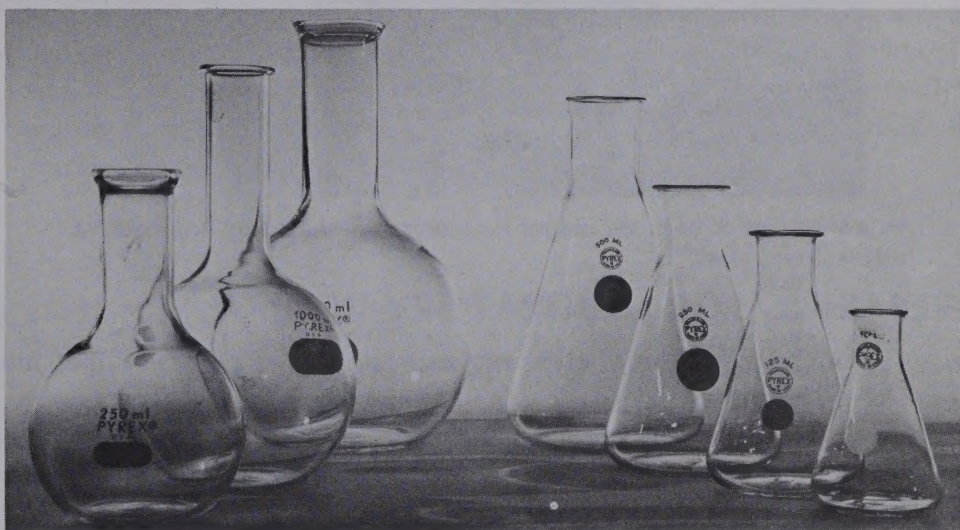


An assortment of cork (above) and rubber (below) stoppers. Cork stoppers are usually available only as solid stoppers. If holes are needed, they must be bored by the user. Rubber stoppers, by contrast, can be purchased as solid, one-hole, or two-hole stoppers. The sizes of both cork and rubber stoppers are designated by a number scheme.

KEYS TO THE CHEMISTRY LABORATORY

THE IMPORTANCE OF LABORATORY WORK TO CHEMISTRY

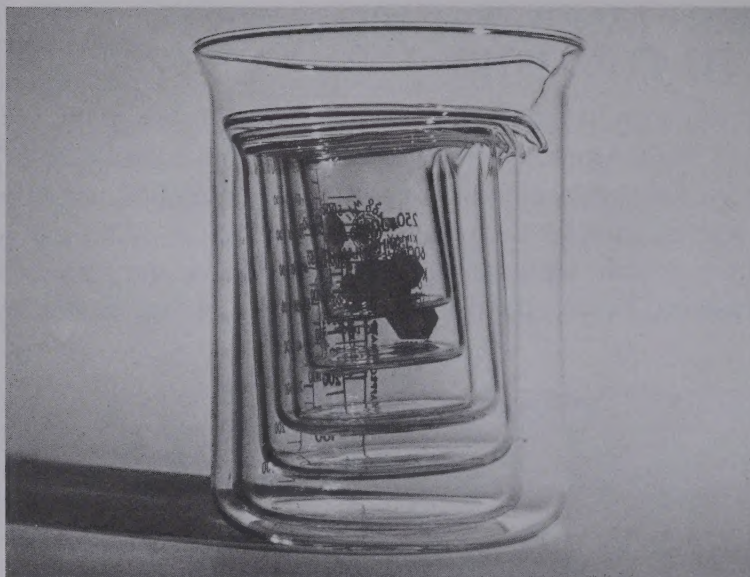
Your laboratory work is the foundation upon which much of the text, *Keys to Chemistry*, is built. More than half your time in this course will be spent in the laboratory. Laboratory work is essential to chemistry because experimentation is the starting point for chemical information. Scientists make observations in the laboratory, then formulate principles, concepts, and theories based on their observations. Finally, the theories are refined through laboratory tests.



An assortment of the various sizes of two kinds of laboratory flasks: left, Florence flasks; right, Erlenmeyer flasks. Both kinds are made of heat-resistant glass so that liquids can be heated in them over a gas burner or an electric hot plate.

As a student of chemistry you will be a scientist. You will do a wide variety of laboratory activities designed to help you make discoveries that will be new to you. The laboratory investigations you will carry on during your study are planned to (1) acquaint you with common laboratory equipment and apparatus, (2) familiarize you with experimental techniques and procedures, (3) provide data which can serve as a basis for discussion, and (4) answer questions which may arise from assigned activities and discussion. Some experiments are designed to let you gather facts for yourself. Others are not experiments in the usual sense of the word, but exercises to help you discover relationships.

These laboratory activities have been planned very carefully. Their real value will come from the thought you give to them and the care with which you perform them. You must learn to be a good observer, to measure precisely, and to control the conditions under which each experiment is carried out.



An assortment of beakers, nested for storage. Beakers too, are made of heat-resistant glass.

PREPARING FOR LABORATORY WORK

Because you have had little laboratory experience, the early activities are more highly structured and contain explicit directions. As you gain experience, the directions become less and less structured. You will be able to exercise imagination and ingenuity. In addition to the investigations given here, you can design experiments of your own. These may grow out of required activities or out of class discussion or from reading current science journals. Watch for problems that can be solved by experimentation and learn to solve them in this way. This is the essence of science and will be the most rewarding part of your study.

It is very important that you understand the directions and point of each experiment before starting to work on it. Study the background paragraph before each experiment carefully. It introduces you to the reasons for doing the experiment and often provides essential information about how to do the experiment.

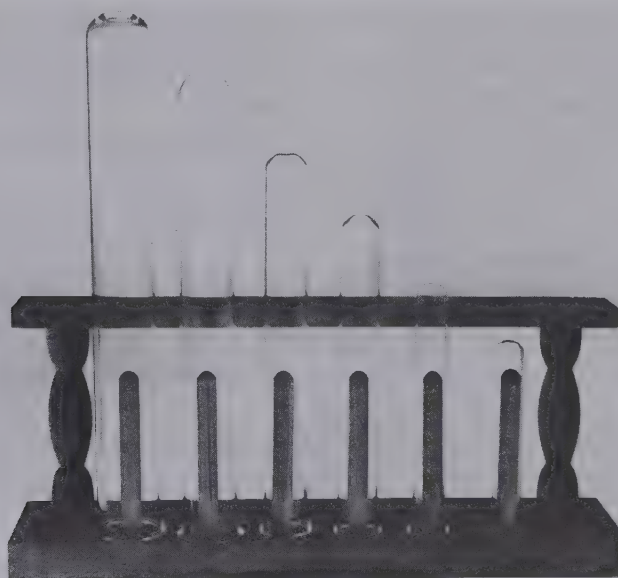
Before starting work on an experiment, it is a good idea to read the procedure through once. As you read, decide what data you will have to record and prepare data tables. Then when you start the experiment, your observations can be recorded systematically, which makes the record meaningful for future use. By studying each experiment beforehand, you will know what you are doing. The experiments will be much more interesting and will seldom need to be done over.

Study the safety precautions and the general rules for conduct in the laboratory given in this introductory section. Work to develop good technique in using laboratory equipment. Take pride in doing careful, precise work and in leaving the lab bench as clean as you found it.

You will gain independence and self-reliance by working alone when possible. You will also gain confidence in your ability to make decisions and to judge the results of your experimental work. Don't rely on someone else and don't let anyone else use you as a crutch. Approach your work with an open mind. Condition yourself to look at your results from many different viewpoints. Use your imagination and common sense.

Mathematical calculations and chemical equations are used only when needed in the experiments. Because they are used in a very practical context you will be able to do them and understand them.

Once you have learned many of the techniques of chemical experimentation, use your knowledge to design your own experiments. To do so can be one of the most rewarding experiences in chemistry. However, for your safety and that of others, be sure your teacher approves such experiments before you try them in the laboratory. Chemistry happens to be an area where what you don't know can hurt you.



A wooden test tube rack and an array of test tubes of various common sizes.

RULES OF CONDUCT IN THE LABORATORY

Certain rules of conduct, listed below, are advisable in a science laboratory. Study them and then list a reason for each rule in your laboratory notebook.

1. Always maintain a businesslike attitude.
2. Do *not* engage in practical jokes.
3. Loud talking and excessive discussion of matters unrelated to the investigation under way is unwise.
4. Dispose of waste chemicals as indicated by the instructor. This is especially important when flammable materials are used.

5. Never return unused reagents to the stock bottles.
6. Use *only* the amount of each reagent called for in the experiment.
7. Wash your hands thoroughly after a laboratory period.
8. Always leave your workbench clean and dry.
9. Be *sure* the gas and water outlets are turned off completely after use.
10. Any time you do not understand directions, ask the instructor for help.

SAFETY IN THE CHEMISTRY LABORATORY

Your personal safety and that of others working near you depends upon the care with which you observe the rules listed below. Study these rules and follow them at *all* times. Since you are more likely to observe a rule if you know the reason for it, list the reason for each of the following safety regulations in your laboratory record book.

1. Know where the fire extinguisher is and how to use it.
2. Know the location of the safety shower and the eyewash fountain and how to use them.
3. Know the location of the fire blanket and how to use it.
4. Always wear eye protectors when doing an experiment.
5. Never wear contact lenses in the laboratory.
6. Perform *only* those experiments in the laboratory manual or ones that have been approved by your instructor.
7. Never taste chemicals unless instructed to do so.
8. Never leave flames unattended.
9. Never use a flammable liquid near an open fire.
10. Never pour a flammable liquid into the sink.
11. Notify your instructor of any accident, no matter how minor it may seem to you.
12. Check the label on all reagent bottles *twice* before using.
13. If spillage of an acid or a base occurs, flood the area with water immediately and notify your instructor.
14. When diluting acids, *always* pour the acid into the water. Remember, A to W!
15. When inserting glass tubing, glass rod, or a thermometer into a rubber stopper or rubber tubing, always protect your hand with several thicknesses of cloth. Lubricate the glass with water or vaseline before inserting it into the stopper.
16. When heating the contents of a test tube, keep it *tilted* and moving in the flame with the mouth pointed away from yourself and your neighbors.
17. Wear a protective apron at all times.
18. Do not wear long, loose sleeves or a loose coat in the laboratory.
19. If you have long hair, tie it back with a rubber band or ribbon while working in the laboratory.
20. Bracelets and dangling jewelry should always be removed while working in the laboratory.

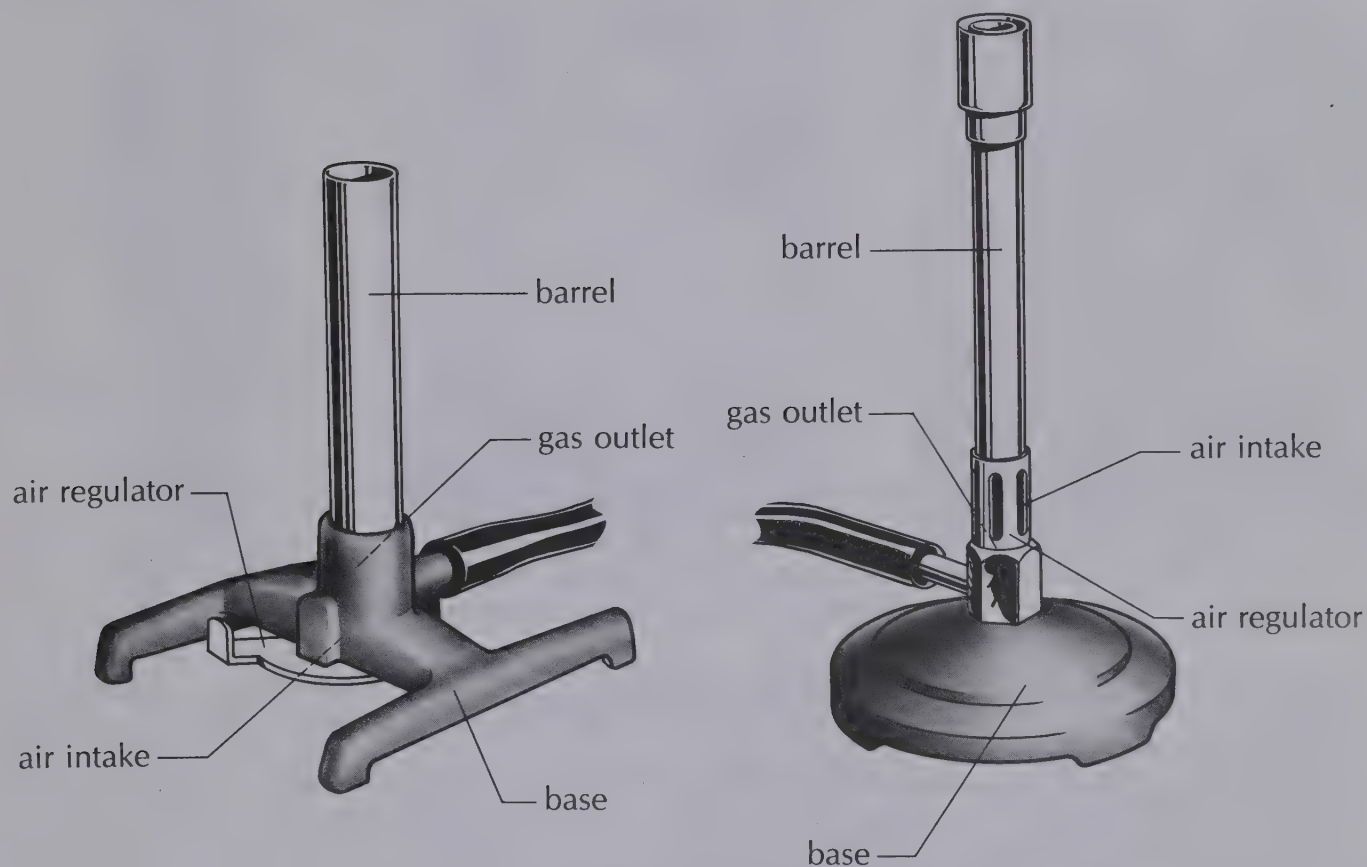
LABORATORY TECHNIQUES AND PROCEDURES

THE GAS BURNER

Since you will be using the gas burner more than any other single piece of equipment, it is important that you learn how to use it most effectively.

Parts of the Burner

There are several types of gas burners in common use. Although differing somewhat in construction, they are quite similar in operation. Examine your particular burner and identify its parts.



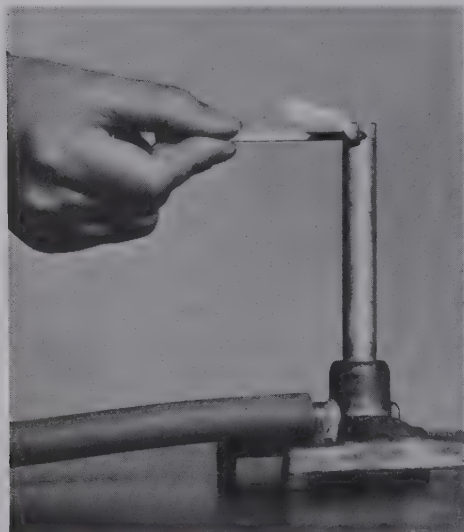
Two types of gas burners commonly used in chemistry laboratories.

1. *Base.* The base is the support for the burner.
2. *Barrel.* The barrel is a chamber in which the fuel gas and air can mix. Some burners are constructed so that the air intake can be regulated by turning the barrel. Others have a collar near the base which can be turned to let in more or less air as required to provide the type of flame you want. Practice regulating the flow of air in your burner.
3. *Gas adjustment screw.* Some burners have a screw in the base to regulate the amount of gas entering the barrel. Others do not have this screw. The gas must then be regulated by the outlet to which the hose is attached. Find out how to regulate the flow of gas in your burner.

Lighting the Burner

CAUTION: Do not lean over the burner so that your face is close to it while lighting it. Instead be a comfortable arm's length from the burner.

1. Close the air vent before lighting the burner. The best way to light the burner with a match is to bring the latter alongside the top of the barrel. This stops a rush of gas-air mixture from putting out the match before the burner lights. If a striker is used to light the burner, hold the flint and steel over the burner and rub the flint firmly across the steel. You may need to use both hands to assure a strong spark.

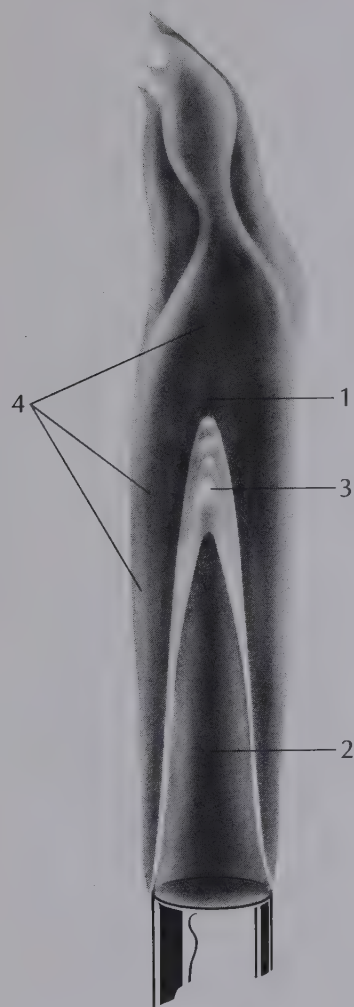


(Left) Why bring the match alongside the burner instead of down on top of the barrel when lighting it? (Right) When lighting a burner with a striker, one strong spark is what is needed. To get the one strong spark, the flint must rub very firmly against the steel.

2. Once lighted, adjust the gas and air flow of your burner so that the flame is about 5 cm tall. This is high enough for most laboratory work and is safer than a taller flame.
3. Occasionally, the flame may flash back and begin to burn at the base of the burner. This is known as striking back. If this happens, turn off the gas and let the burner cool. Be careful not to touch the barrel; it may be quite hot. Once cool, relight the burner.
4. If a gap exists between the top of the barrel and the flame, decrease the air supply.
5. If the flame makes a roaring noise, decrease the air supply.

The Burner Flame

1. Light the burner and adjust it so that you have a blue, almost invisible flame. This is the type of flame you will normally use.
2. Examine the flame closely. Note the different cones that are visible.
 - a. *Outer cone.* This is the hottest part of the flame because fuel combustion is most complete here.

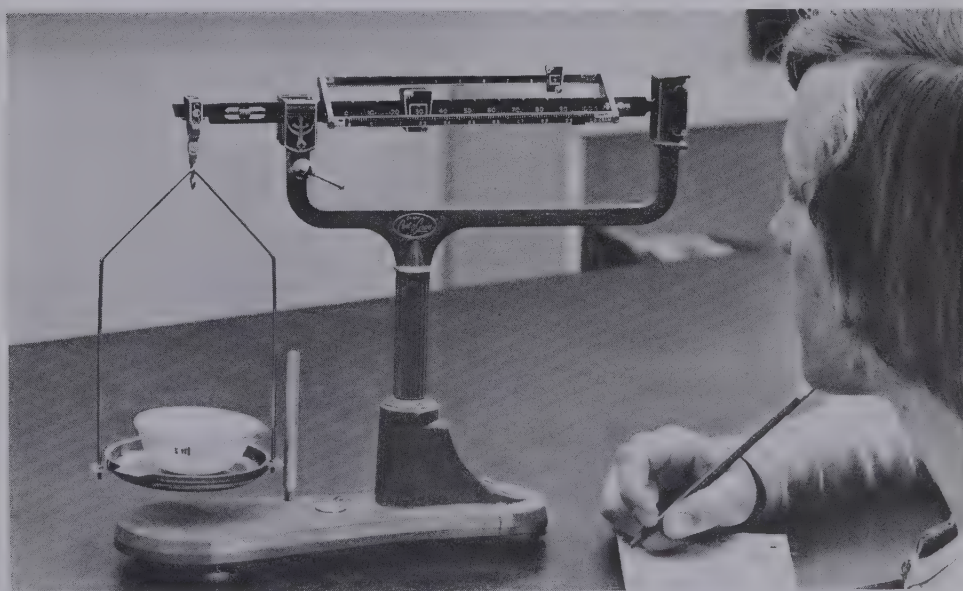


- b. *Inner cone.* This cone is blue in color and not as hot as the outer cone. The gas has not had a chance to burn completely in this region.
 - c. *Base cone.* There is a dark cone in the interior of the flame. It is the mixture of gases before they start to burn.
 - d. Occasionally you will need to use a yellow flame. This is produced by shutting off most of the air supply. With little air in the fuel mixture, the gas burns incompletely. Unburned carbon in the flame glows to produce the yellow color. This type of flame deposits soot on objects placed in it. Do not use a yellow flame unless instructed to do so.
3. Adjust the flame of the burner so that it is about 8 cm tall. Place a piece of cardboard in the flame with one edge resting on the top of the barrel of the burner. Remove the cardboard from the flame before it ignites, but not before a scorch pattern is formed on it. Note the shape of the scorch pattern. What does this indicate about the relative temperatures of the different cones?
 4. Turn off the burner. Thrust a pin through a match just below the head. Suspend the match in the top of the barrel of the burner. Be sure the match head is in the center of the barrel. Now light the burner. Note what happens to the match head. What does this suggest about the dark, inner portion of the flame?

MEASURING WEIGHT

The triple beam balance is recommended for most of your laboratory work because it can distinguish differences in weight of 0.01 gram. Since types of triple beam balances differ, your instructor will demonstrate the type available to you. There are, however, a few directions which apply to all balances.

A burner flame is the site of a complex mixture of chemical events which you may wish to investigate as you acquire a working knowledge of chemistry. For the present you should know that (1) in the diagram is the hottest point in the flame, (2) is in a zone composed of unburned gas and air, (3) is in a zone of burning gases, and (4) is in a zone of partially burned, still hot, gases.



A triple-beam, single-pan balance.

1. Balances are sensitive instruments and must be treated carefully if they are to continue to work properly. Do not try to weigh an object on the balance if you think it is beyond the capacity of the balance. So, determine the capacity of the balance you will use.
2. Do not place chemicals in a metal weighing pan without first protecting the metal with a piece of plastic or lightweight paper. Many chemicals will corrode metal and cause deterioration of the pan. Placing a chemical directly in the metal pan may also contaminate the next chemical weighed.
3. Before using any balance, make sure that all rider weights are in their proper place and that the pointer swings freely. You need not wait for the pointer to come to rest on the zero point, although it should oscillate the same distance on either side of the zero point.
4. Do not weigh hot objects; they may damage the balance. In addition, convection currents caused by the heat may cause inaccurate readings.
5. Balances should be kept dry. If you accidentally spill a liquid on a balance, rinse it off and dry it thoroughly.

MEASURING LIQUID VOLUME

In chemistry you will use several different types of apparatus to measure the volume of liquids. These include the graduated cylinder, the buret, the volumetric flask, and the calibrated pipet. Some of these are designed to deliver a certain volume of liquid. Some are designed to contain a certain volume of liquid. The letters *TD* on the apparatus means "to deliver"; the letters *TC*, "to contain." Usually the temperature at which the calibration is made is specified on the apparatus. The calibration temperature is normally 20°C, or about room temperature.



An assortment of graduated cylinders ranging from a capacity of 1000 cc to 10 cc.



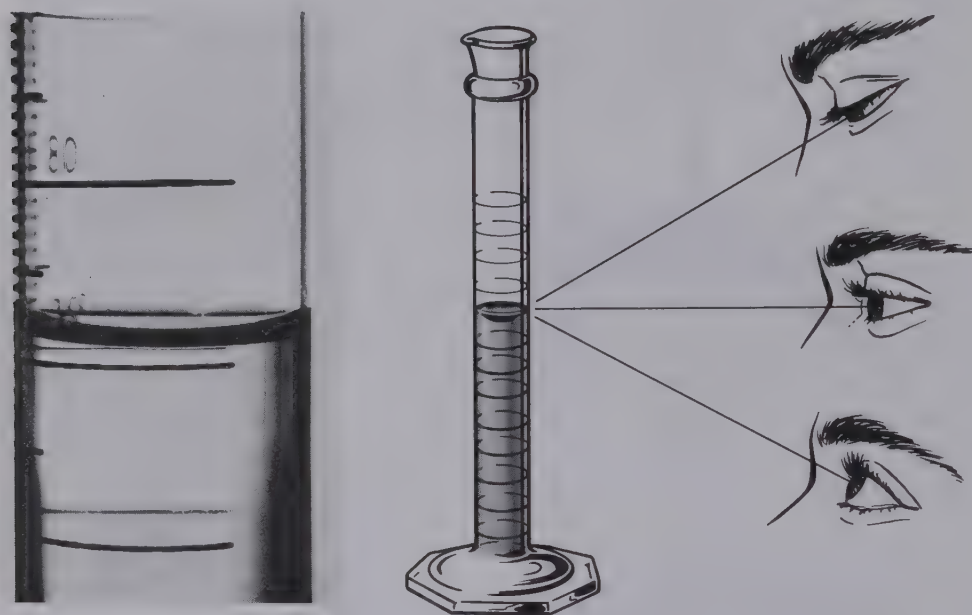
(1) Support stand, (2) buret clamp, (3) 50 cc graduated burets (note that the scales on the burets increase from top to bottom so that they conveniently measure what is removed rather than what they contain), (4) 10-cc graduated pipet, (5) graduated cylinder, (6) ungraduated or uncalibrated cylinder, (7) 250-cc Erlenmeyer flask.

Calibration

Examine the graduated cylinder you will be using. Determine the volume indicated by each of the graduation marks. Some cylinders measure in milliliters (ml), some in 0.2 ml, and others in 0.1 ml. Always use the cylinder with a scale most appropriate to the accuracy you will need.

The Meniscus

The surface of liquids when viewed in glass cylinders is always curved. Most of the liquids you will be measuring will have a concave meniscus. Certain precautions must be noted if your volume readings are to be consistent and precise.



A closeup view of a meniscus; note the curvature in the liquid surface. How will the position of the eye in relation to the liquid level affect the reading of the volume?

1. The meniscus must be viewed along a horizontal line of sight. Do not try to make a reading when looking up or down at the meniscus. Hold the apparatus up so that its sides are at a right angle to your eye.
2. Always read a concave meniscus from the bottom. This gives the most precise measurement of volume, because the liquid tends to creep up the sides of a glass container. Many plastic cylinders do not form a meniscus. If you are using a plastic graduate and no meniscus is noticeable, read the volume from the level of the liquid.

TRANSFERRING SOLIDS FROM REAGENT BOTTLES

Many experiments require you to transfer a certain amount of solid from a stock, or reagent, bottle. If many students are to use a reagent from a common

container, it is very important that it not be contaminated. Some general rules to be observed in transferring a solid from a stock bottle are:

1. Under *no* circumstances should a solid be returned to the stock bottle. (Most solids used in beginning chemistry are white. A student might accidentally replace a solid in the wrong stock bottle. This would contaminate the bottle for others and in some cases might produce a dangerous mixture of chemicals.)

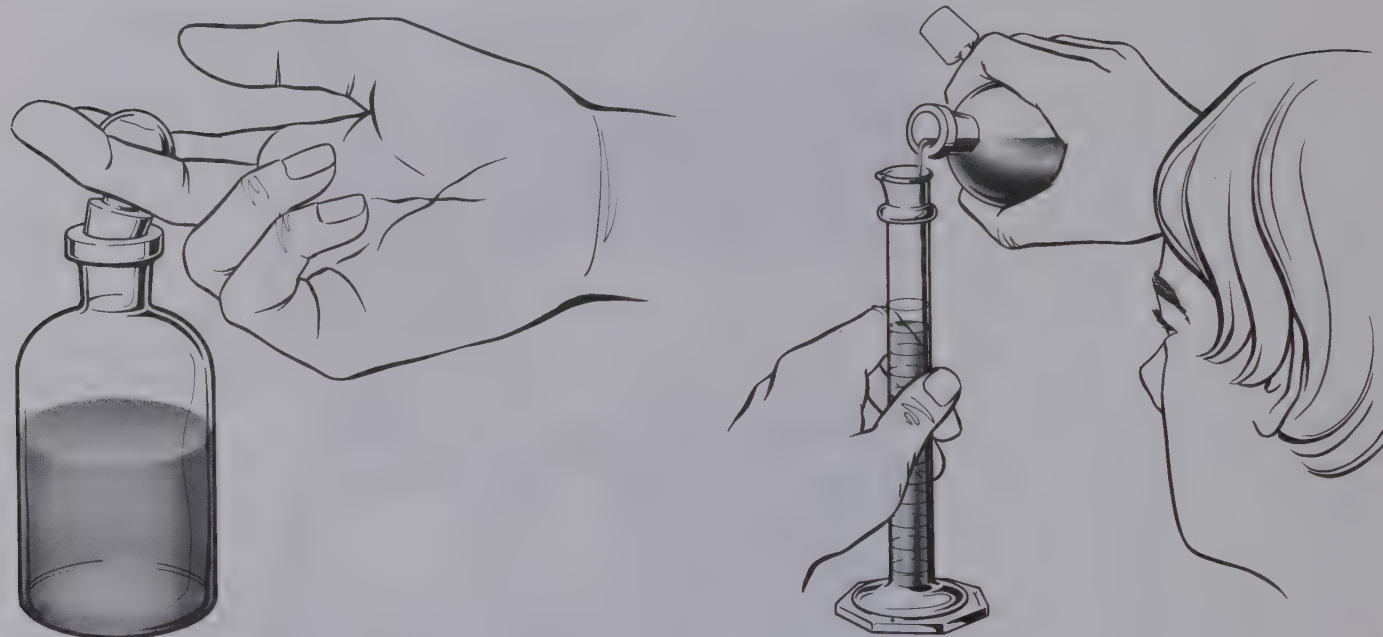


Using a spatula to transfer a small amount of solid reagent from a stock bottle to a beaker.

2. Most solids are required in small quantities. Try to estimate the quantity of a solid that you need. Remove the stopper or cap from the stock bottle. If it cannot be held in the hand conveniently, place the cap on the desk with the *top down*. This prevents contamination. Use a *clean* spatula to remove the amount of solid you need.
3. If the quantity of solid is very small, it may be transferred directly from the spatula to a beaker or test tube. If the quantity is too large for this method, place the solid on a small piece of creased paper. The paper can act as a trough to pour the solid into the beaker or test tube without spilling.
4. Since many reagents react with the moisture or the oxygen in air, be sure to cap or stopper the bottle as soon as possible after removing the solid.

TRANSFERRING LIQUIDS FROM REAGENT BOTTLES

When transferring liquids from a reagent bottle it is very important that the stopper not be contaminated. Many reagent bottles have what is called a pennyhead stopper, which can be held conveniently while pouring liquid. Do *not* lay this stopper on the table; it can pick up dirt. There is also danger of damaging the desk top if the liquid is corrosive.



Why should you hold the stopper as well as the bottle while pouring from the bottle?

General rules for transferring a liquid to a test tube or graduated cylinder are:

1. Remove the stopper and hold it between your fingers as shown.
2. Hold the test tube or graduated cylinder at *eye level* and pour the liquid *slowly* until the desired volume has been transferred.
3. Replace the stopper in the reagent bottle.
4. If any of the liquid runs down the outside of the bottle, rinse it with water before returning it to the shelf. This is particularly important if the liquid is corrosive. If the bottle is not rinsed, it may damage the shelf. Also, the next student who uses the bottle may be burned by the liquid. Remember: *you might be that student!*
5. Since many experiments do not require a precise volume of liquid, it is convenient and timesaving to estimate the volume of liquid which equals 1 or 5 cc. To do this, measure 1 ml of water in a graduated cylinder and pour it into a test tube of the size you will use most frequently. Note the height of the water in the test tube. This will enable you to estimate volume when an experiment calls for approximately 1 cc of any liquid.

- a. Repeat this procedure, using 5 ml of water. This technique will save you many minutes in the course of a term. You will also save having to wash your graduated cylinder unnecessarily.
- b. With the pipet you will normally use, determine the number of drops of water equal to 1 cc. Although the drops may vary a bit in size, this technique may also save time.

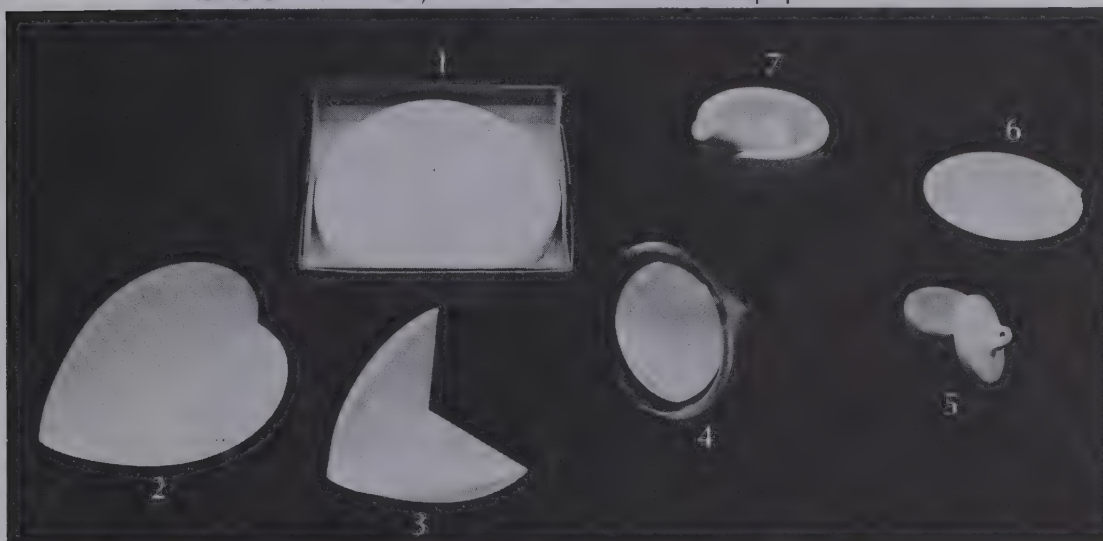
FILTERING PRECIPITATES

In many experiments you will need to filter a mixture to separate solid and liquid material. The term *precipitate* refers to the solid that forms when two solutions are mixed. Filtering is accomplished by pouring the mixture of solid and liquid into a filter paper cone in a funnel. The solid is caught by the filter paper. The liquid, often called a filtrate, passes through the filter paper to another container.

Folding Filter Paper

Filter paper must be folded so that it fits into the funnel you intend to use. If you have ever folded a sheet of paper to make a drinking cup, you already know how to fold a filter paper. In case you are not sure, follow these directions.

1. Fold the paper in half.
2. Fold the paper once again, so that it is quartered.
3. Place the folded filter paper into a funnel so that three thicknesses of the paper are on one side of the funnel and only one side is against the opposite side of the funnel.
4. Moisten the filter paper with water after placing it in the funnel, to prevent its popping out of the funnel. NOTE: If the liquid to be filtered is *not* miscible with water, *do not* moisten the filter paper with water.



(1) Box of filter paper, (2), (3) stages of folding a piece of filter paper, (4) the folded filter placed in a funnel, (5) crucible and cover, (6) evaporating dish, (7) mortar and pestle.

Filtering Procedure

1. If the amount of liquid to be filtered is small, place the funnel in a test tube supported in your test tube rack.
2. If the amount of liquid is too large for your test tube, place the funnel in a support and with a beaker under it so that the stem of the funnel touches the side of the beaker.

Washing the Residue

In many cases you will be instructed to "wash the residue." This is important. It removes any residual solution which clings to the residue. Follow this procedure:

1. If you need to work with the filtrate, remove the funnel from the test tube or beaker containing the filtrate and place it over another test tube. Using a plastic squeeze bottle, spray 2–5 cc of distilled water into the funnel to wash the residue. Discard this filtrate.



Using a wash bottle to wash down a piece of filter paper after filtering is completed.

2. If you do *not* need to work with the filtrate, you need not remove the funnel from the original test tube or beaker before washing the residue.

Drying the Residue

In many cases the residue must be dried for further study. In this case, remove the filter paper from the funnel and unfold it. Wet filter paper is fragile, so do this operation very carefully so as to avoid losing any part of the residue. Dry the residue as indicated by your teacher.

Centrifuging

If a centrifuge is available, separation of a precipitate from a liquid can be done much faster than by filtering. A centrifuge may supply a force up to 1000 times that of gravity, and only a few seconds are required for the operation.



A centrifuge.

Your instructor will demonstrate the proper use of a centrifuge. The following precautions apply to the use of any centrifuge.

1. If you are centrifuging only one test tube of material, *always* counter-balance it with another test tube containing an equal amount of water placed *opposite* the one containing the experimental material.
2. The precipitate will be forced to the bottom of the test tube. The liquid can then be poured off or removed with a dropper.
3. Occasionally a test tube breaks while being centrifuged. If this happens, notify your instructor at once so that the liquid can be cleaned out of the centrifuge before corrosion occurs.

RECOVERING A SOLID FROM A LIQUID

The solubility of solids varies greatly. When a solid dissolves in a liquid, it can be recovered by evaporating the liquid. The solid will remain as a residue. This can be accomplished in two ways: by the drop method and by the quantity method. If the amount of dissolved solid is not important and all you need to know is *whether* solid is dissolved in the liquid, the drop method is used.

Place one or two drops of the solution on a glass slide. Hold the slide between the thumb and forefinger and move it back and forth above the flame of a gas burner until the liquid evaporates. It is important that the glass slide *not* be placed directly in the flame. Excess heat will break the slide and



Drying a small amount of solution on a microscope slide.

may decompose the residue. Therefore, hold the slide at least 2 cm above the tip of the flame. Holding the slide in your hand will help guarantee that it will not become too hot.

If the residue from evaporation is to be used in further analysis, the quantity method is used.

1. Place the solution to be evaporated in an evaporating dish.
2. Set up a ringstand and ring with the ring about 5–6 cm above the barrel of a gas burner.
3. Place an asbestos mat on the ring. Set the evaporating dish of liquid on the mat.
4. Light the gas burner and adjust the flame to a height of about 5 cm.



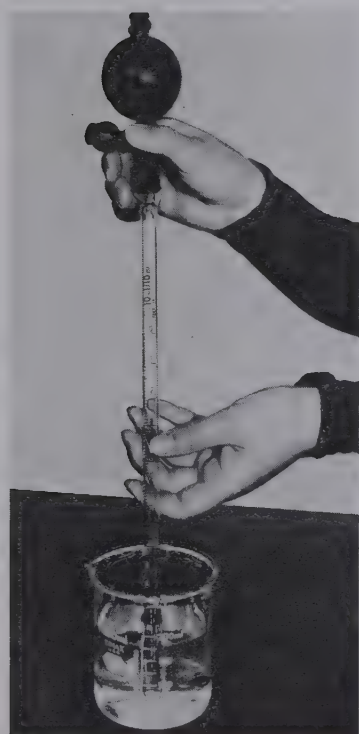
Arrangement for evaporating a larger amount of solution.

5. Heat the dish until *most* of the liquid is evaporated. Then turn off the gas and let the residual heat finish the evaporation process. If the flame is left on until the liquid has completely evaporated, spattering may occur and some of the residue will be lost.

PIPETTING A LIQUID

The dropper pipet is normally used to transfer small quantities of a liquid to a test tube. It is similar to a medicine dropper except that it has no rubber bulb (acids and bases corrode rubber). Never insert the pipet into a stock bottle of reagent. If you do, you will contaminate the reagent. Always pour a small quantity of the reagent into a test tube first, then use the pipet to take liquid from it. With a little practice, you can release the liquid from the pipet, drop by drop, into another container.

Occasionally you may find it convenient to use a calibrated pipet. This pipet has graduation marks like a graduated cylinder. The liquid should be drawn into the pipet to a point just above the calibration mark and held there by placing a finger over the top of the pipet. The liquid level can then be adjusted by carefully admitting air to the top. In pipetting corrosive liquids a suction bulb or a source of vacuum should be used to draw the liquid into the pipet.

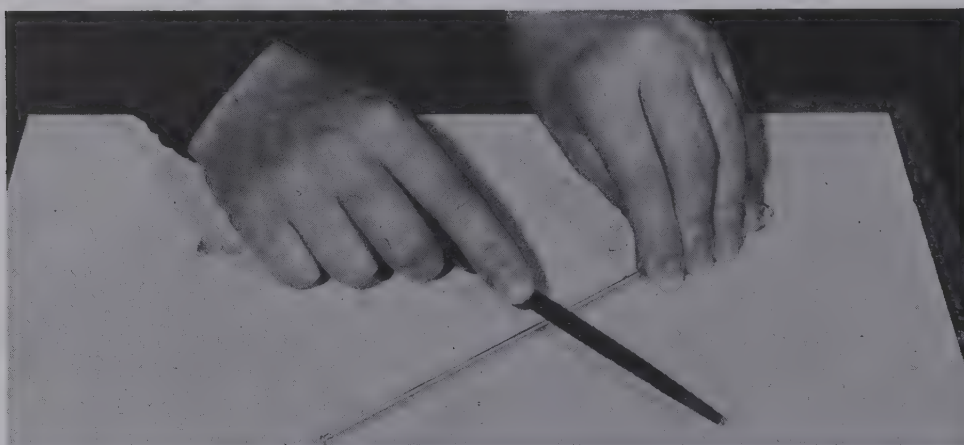


Using a pipet equipped with a pipetting bulb to measure a small quantity of liquid.

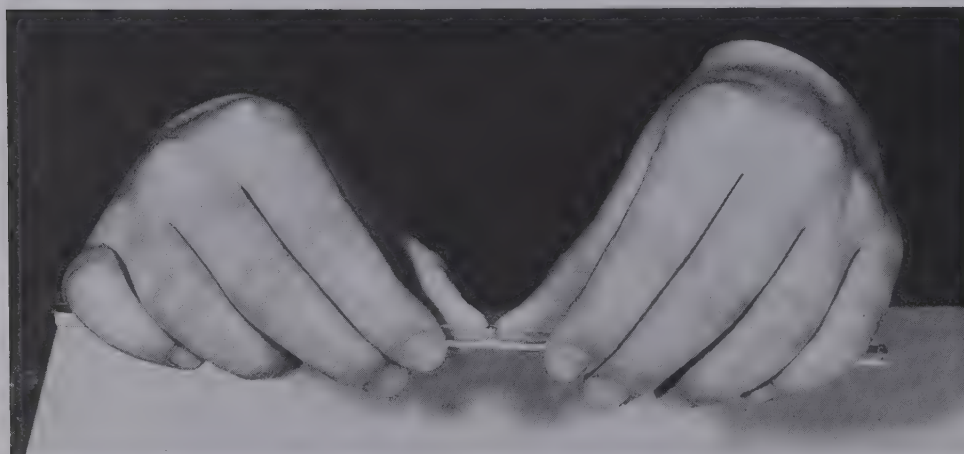
WORKING WITH GLASS

Cutting Glass Tubing and Glass Rod

1. Place the tubing or rod flat on your desk. Hold it firmly with one hand.
2. Use a sharp, triangular file to make a heavy scratch across the glass where it is to be cut. It is unnecessary to "saw" a deep scratch into the glass. One good stroke is usually sufficient.
3. Grasp the glass, with the scratch facing away from you. Place your thumbs opposite the scratch and bend the ends of the glass toward your body. Fire-polish the glass before using it.



Scratching a piece of glass tubing as the first step in breaking it in two.



Holding the scratched piece of glass to break it.

Fire-Polishing Glass Tubing and Glass Rod

1. Light the gas burner; adjust it to a blue flame.
2. Rotate the glass rod or tubing with the end in the hot part of the flame. Heat it just enough to get rid of sharp edges. Do not heat the glass tubing until the opening is reduced in size.

3. Place the hot glass on an asbestos mat to cool. **CAUTION:** Glass cools slowly! Keep in mind that hot glass looks no different from cold glass. Keep track of what part you heated and never pick up someone else's pieces of glass.

Inserting Glass into Rubber Stoppers and Rubber Tubing

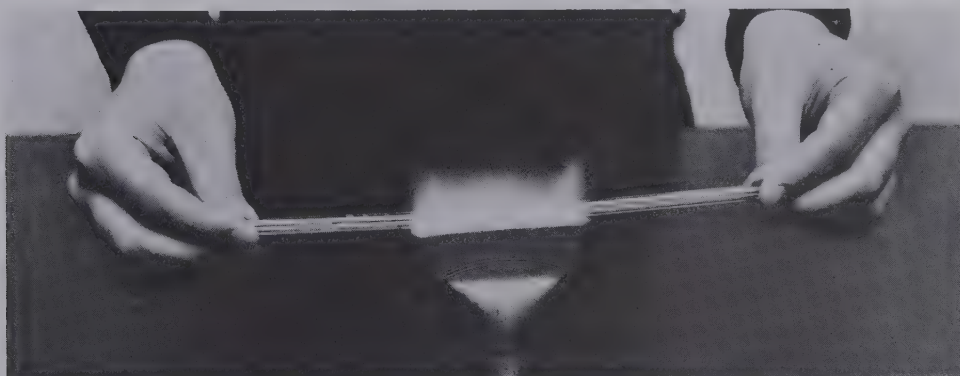
When inserting a thermometer, glass tubing, or glass rod into a rubber stopper or rubber tubing, four rules should be followed to avoid injury.

1. Always fire-polish the ends of a piece of glass rod or tubing before inserting it in a stopper or rubber tubing. Rough edges on the glass makes insertion more difficult and may damage the rubber.
2. Always wet the glass and the hole into which it is to be inserted with water before trying to insert the glass.
3. Apply only gentle pressure to get the glass inserted. If excess pressure is needed, you have a mismatch. A hole in a stopper can be enlarged; and a larger size of rubber tubing can be used.
4. Protect your hands with several thicknesses of cloth. If the glass breaks while you are forcing it into the stopper, you will be less likely to injure your hands.

If left in a rubber stopper for any length of time, glass becomes "frozen" and is hard to remove. If this happens, use either a sharp knife to slit the stopper or a cork borer that just fits over the glass to expand the hole in the stopper.

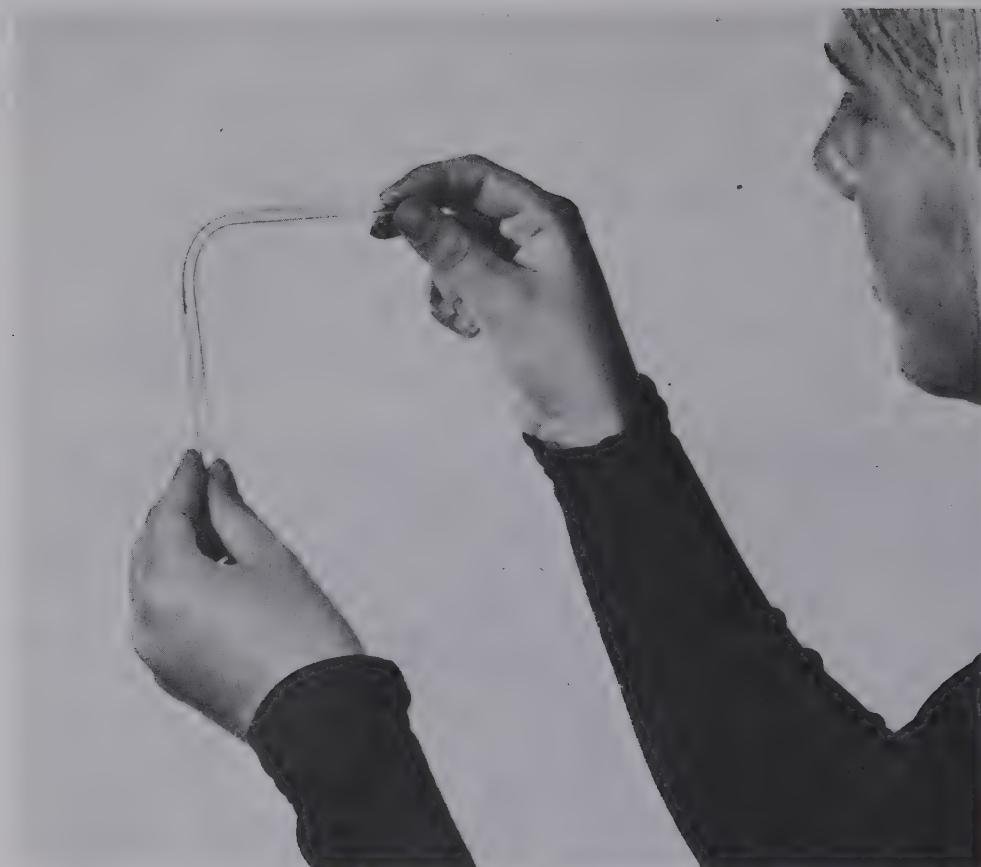
Bending Glass Tubing

1. Fit your gas burner with a wing top. This gives a wide flame that will heat a long portion of glass tubing.
2. Light the burner. Adjust to a blue flame about 5 cm high.



Heating a piece of glass tubing before bending it.

3. Rotate the glass continuously in the flame so that it is heated evenly on all sides. Glass contains some sodium and will cause the flame to become yellow when it is almost hot enough to bend. Continue heating until the glass begins to sag.



Inspecting the completed bend.

4. When the glass is hot enough to bend easily, remove it from the flame and gently, but quickly, bend it to the desired angle. Try to produce a smooth curve rather than a sharp bend. When bending, try to keep the two ends in the same plane.
5. Place the hot glass on an asbestos mat to cool. If the bend is in a plane it will lie flat.

1-1 • PRACTICE IN SCIENTIFIC OBSERVATION

Good observations are the key to success in science. They also make daily living much more interesting. Attention to details, alertness to change, and awareness of the factors which cause change are important in making good observations. A good observer needs both practice and patience.

Observations may be of two types: qualitative and quantitative. Qualitative observations are general and usually expressed in words. An example of a qualitative observation is, "It was hot most of the day." One might ask, "How hot is hot?" "What is meant by 'most of the day'?" How might this observation be made more meaningful?

Quantitative observations are much more useful than qualitative ones. Quantitative observations are measurements, expressed as numbers. They tell how much, how long, how fast, etc. We could make the above observation quantitative by saying, "The temperature was in the high 90's until 3:00 P.M., when a cloud formed and the temperature dropped into the low 80's."

PURPOSE:

To make and record meaningful observations while watching materials change.

PROCEDURE:

Part I: Observing Sugar Cubes Dissolve in Water

1. Fill a 250 cc beaker $\frac{3}{4}$ full of tap water. Place it where you can observe it from all angles without disturbing it.
2. Drop two sugar cubes into the water. Record your observations for 15 minutes.



3. At the end of 15 minutes carefully lift the beaker. Observe the contents as you gently tilt the beaker to and fro. Record additional observations.
4. Empty the beaker in the sink. Wash the beaker and return it to your locker.

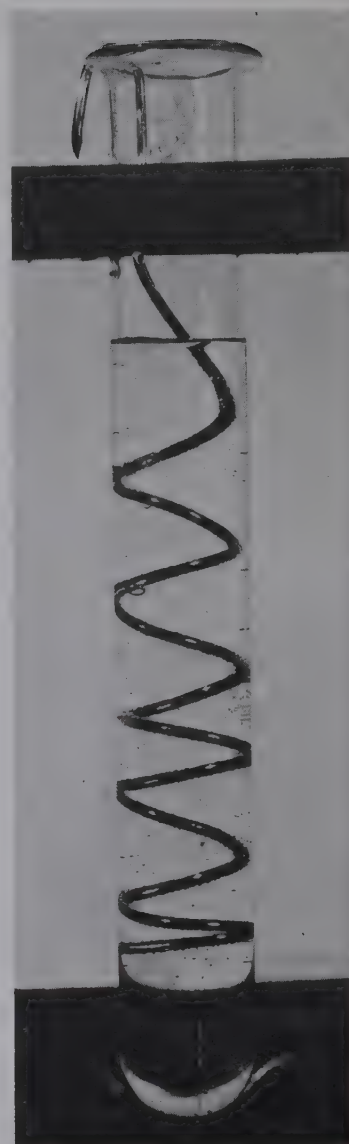
Part II: Observing a More Dramatic Change

1. Obtain a 100 mm \times 13 mm test tube and a piece of copper wire about 20 cm long. (Note: The wire must not be coated with shellac or any other material.)
2. Coil the copper wire around a pencil, leaving a tail to hook over the top of the test tube. Be sure the coil is fairly open. The bottom end of the coil should be 2–3 cm off the bottom of the test tube.
3. Fill the test tube $\frac{3}{4}$ full with silver nitrate solution. (CAUTION: Silver nitrate solution will stain your fingers brown. It is not harmful, although the stains are not very pretty. If you do spill some, wash it off immediately to minimize the stain.)
4. Place the copper coil in the solution; support it by hooking the tail over the top of the test tube.
5. Observe the contents of the tube for 20–30 minutes. Record both qualitative and quantitative observations.
6. You may wish to leave the test tube overnight so you can make further observations the next day. Consult your teacher about how and where to store the tube.

REPORTING YOUR OBSERVATIONS

Make a formal report of your investigations, dividing it into Part I and Part II. Include the following in each part:

- a. qualitative observations
- b. quantitative observations
- c. inferences or interpretations for some of the observations
- d. hypotheses (possible explanations) for some of your observations
- e. plans for a further investigation to test at least one of your hypotheses.



1-2. SOME USEFUL LABORATORY TECHNIQUES

In laboratory work it is often unnecessary to measure the precise volume of a liquid to be used. When this is true, it is convenient to be able to estimate 1 cc, or 5 cc, or whatever, in test tubes, beakers, etc. This saves time and trouble in washing graduated cylinders, which must be used if a precise volume of liquid is called for. In this experiment you will learn how to estimate volumes of liquids in common apparatus.

Hydrogen, oxygen, and carbon dioxide are gases given off in many of your experiments. Learning how to test for these common gases now can save time later on. You will produce small quantities of these gases in this experiment and test to identify each one.

Another useful tool in laboratory work is the flow chart. This is a special kind of outline for a procedure and illustrates the idea that "one picture is worth a thousand words."

PURPOSE:

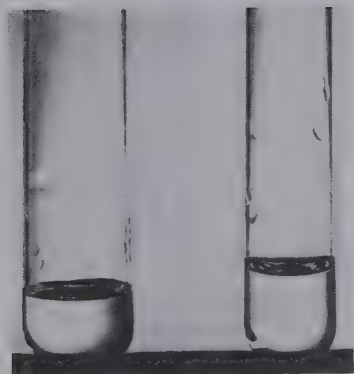
To learn some useful laboratory techniques to make your work easier and more efficient.

PROCEDURE:

Part I: Estimating the Volume of Liquids

A. Number of Drops per cc

1. Review the proper technique for pipetting a liquid (see page 17 in the technique section of this manual).
2. Fill a small test tube about half full of tap water to use for practice.
3. Practice pipetting until you can release the water one drop at a time.
4. Next, count the number of drops it takes to make 1 cc of liquid in a 10 cc graduated cylinder. (Your graduated cylinder may have markings in ml's. This is the same as cc, or cubic centimeter.) Repeat the process several times until you are confident of your results. Record this information in your laboratory record book.



How much water is in each of these test tubes?

B. Estimating the Volume in a Test Tube

1. Hold a 100 mm \times 13 mm test tube at eye level. Place the lower part of your thumb on the test tube at the level you think would indicate 1 cc of liquid.
2. Pour water into the test tube to the level marked by your thumb.
3. Pour this water into a 10 cc graduated cylinder. See how close your estimate came to being 1 cc.
4. Repeat this process for 2 cc, 5 cc, and 10 cc.
5. If your estimates were not satisfactory to you, practice until they are.

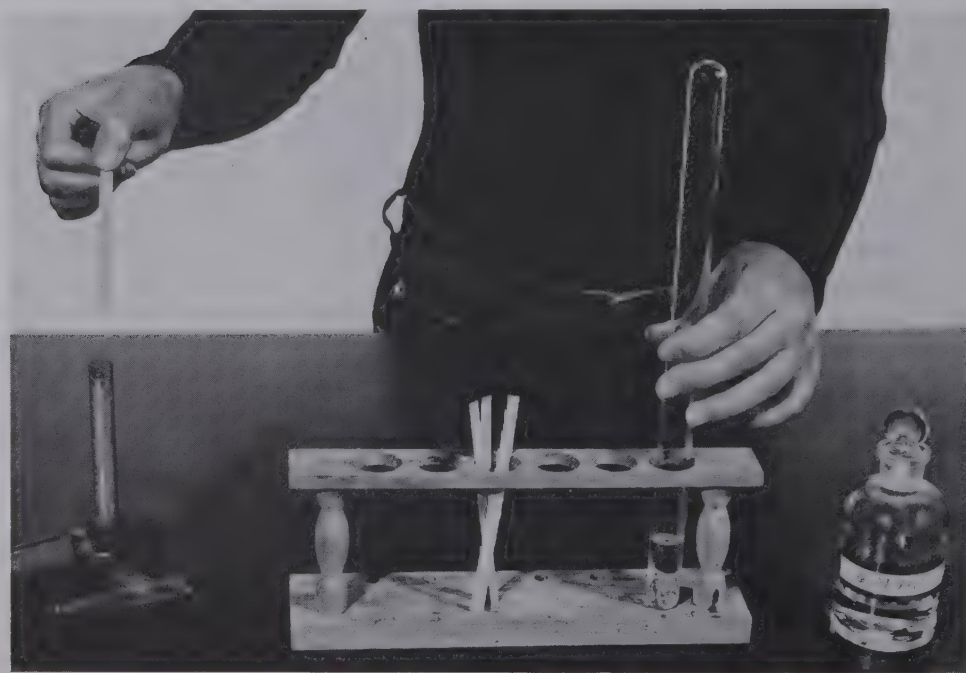
C. Estimating the Volume in Beakers

1. Select a beaker and note the volume the manufacturer has marked on it. For a 50 cc beaker, measure out 50 cc of water in a graduated cylinder. Pour the water into the beaker.
2. Note how much space is left empty. The beaker is calibrated for the "working volume."
3. Now estimate the expected level of liquid in the beaker if you place 10, 25, or 35 cc of liquid in it. Check your predictions by measuring these volumes and marking the various levels with a grease pencil.

Part II: Tests for Common Gases

A. Hydrogen

1. Place a small piece of mossy zinc, Zn, in a 100 mm \times 13 mm test tube.
2. Light the burner and have a wood splint handy.
3. Obtain a second, identical, clean, dry test tube.
4. Carefully pour 2 cc of hydrochloric acid into the tube with the zinc.
5. Place the second test tube over the first. When you think the collecting tube is full of gas, remove it, keeping the mouth down. Quickly bring a burning splint to the mouth of the tube.
6. If the burning splint causes the gas to burn with a "barking" sound, the gas being evolved is hydrogen. You have tested a mixture of hydrogen and air. (CAUTION: Mixtures of hydrogen and air are explosive. Large quantities are dangerous! Do not produce larger quantities of hydrogen in the laboratory without your teacher's permission.)



Technique for collecting a gas.

Technique for testing for hydrogen gas.



B. Oxygen

1. Place 2 cc of a solution of hydrogen peroxide, H_2O_2 , in a 100 mm \times 13 mm test tube. (CAUTION: Be sure to use a 3% solution of hydrogen peroxide.)
2. Light your burner and have a wood splint nearby.
3. Use your spatula to add a tiny pinch of manganese dioxide, MnO_2 , to the hydrogen peroxide.
4. As soon as you notice a rapid evolution of gas, light the wood splint. Then blow it out so that it merely glows, and insert it into the test tube. Note the results.
5. If the glowing splint bursts into flame, the gas evolved is oxygen. This is the test for oxygen.

C. Carbon Dioxide

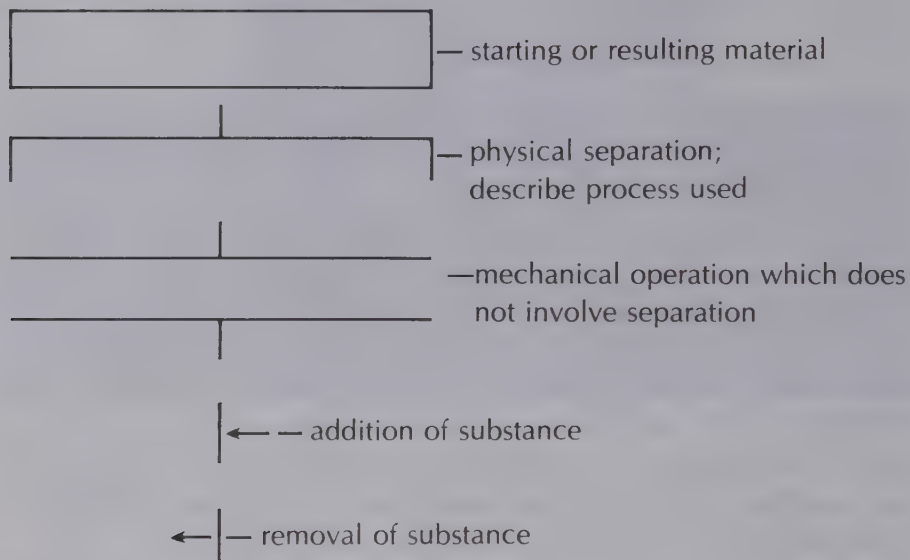
1. Place about $\frac{1}{2}$ cc of baking soda in a 100 mm \times 13 mm test tube.
2. Have a clean stirring rod nearby.
3. Pour about 3 cc of limewater, $\text{Ca}(\text{OH})_2$, in a second test tube.
4. Carefully add about 2 cc of hydrochloric acid, HCl , to the baking soda.
5. As gas is evolved, pick up a drop of limewater with the stirring rod and bring it into the gas. Note the condition of the limewater drop both before and after it has been in the gas.
6. If the drop of limewater turns milky, the gas being evolved is carbon dioxide. This is the test for carbon dioxide. (Note: If enough carbon dioxide is produced in an experiment, the gas may be bubbled through limewater. However, the amount of gas evolved is often too small to use this test.)



Part III: Learning How to Use a Flow Chart

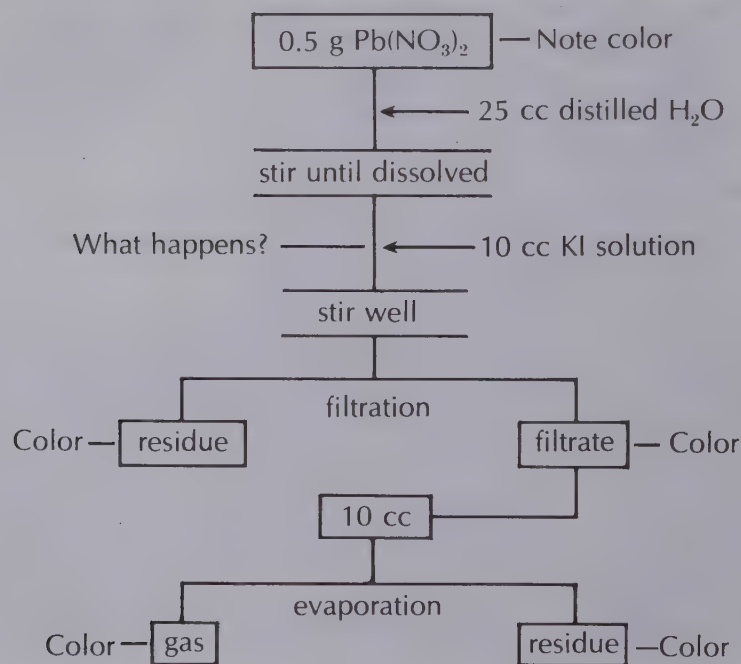
This operation teaches you how to use a flow chart. It also introduces some important laboratory procedures and some terms you need to know. You need not understand all the chemical reactions at this time; that understanding will come later.

Meaning of symbols in a flow chart:



PROCEDURE:

Just follow the directions in the flow chart.



1-3 • LEARNING TO WORK WITH VARIABLES

Many experiments include factors which change in relation to each other. Such factors are called variables. In a controlled experiment, we change one factor and observe its effect upon a second. Often, by graphing the data obtained, we can see a mathematical relationship between the variables that is far more meaningful than looking at a data table. This experiment gives you practice in working with variables, in recording the ordered pairs in tabular form, and in graphing data.

PURPOSE:

To study the relationship between related variables.

TECHNIQUE:

Good data tables are useful when recording data in the laboratory. A sample data table for this experiment might look like this.

Data Table, Experiment 1-3		
Volume of Water	Total Mass (graduate and water)	Mass of Water
0.0 cc	11.2 grams	0.0 grams
1.0	12.3	1.1
etc.	etc.	etc.

Part I: The Relationship between the Volume of Water and Its Mass.



PROCEDURE:

1. Record the mass of an empty 10 cc graduated cylinder.
2. Place 1 cc of water in the graduate. Record the mass.
3. Increase the volume of water, 1 cc at a time, until you have added 10 cc of water. Record the mass of the graduate and water for each cc added.

ANALYSIS OF DATA:

1. Determine the mass of the water used in each of the ten trials.
2. Plot the ordered pairs on a graph. Always plot the independent variable on the horizontal (x) axis and the dependent variable on the vertical (y) axis.)
3. Join the points with a smooth line.
4. Find the slope of the line (rise over run).
5. Write an equation which relates the value of the slope to the two variables.
6. What is the expected mass of water at zero volume?
7. What is the average increase in mass per cc of water?
8. Predict the mass of 100 cc of water.

Part II: The Relationship between Time and Temperature as a Sample of Water Is Cooled from the Boiling Temperature**PROCEDURE:**

1. Prepare your data table before performing this part of the experiment.
2. Fill a 250 cc beaker $\frac{3}{4}$ full of water.



3. Bring the water to a boil.
4. Turn off the burner. Record the temperature of the water.
5. Record the water temperature each minute until the water has cooled to about 35°C .

ANALYSIS OF DATA:

1. Plot a time-temperature graph of the ordered pairs. Again, plot the independent variable on the horizontal (x) axis and the dependent variable on the (y) axis.
2. Join the points with a smooth line.
3. What is the ratio of temperature change to time change during the first minute, the eighth minute, and the last minute of your observation period?
4. Suggest explanations for any change in the ratio of temperature change to time change.

Part III: The Relationship between the Diameter and Circumference of Circular Objects

PROCEDURE:

1. Prepare your data table.
2. Obtain at least five can lids or other circular objects. These should vary from small to large to provide a variety of data.



Two arrangements of circular objects. Is there a constant relationship between diameter and circumference that holds for all of them?

3. Measure the diameter and the circumference of each object. Use significant figures to express the precision of your measurements as you record these ordered pairs in a data table.

ANALYSIS OF DATA:

1. Plot the ordered pairs on a graph. Identify the independent and dependent variables and plot them appropriately.
2. Join the points with a smooth line.
3. Find the ratio of circumference to diameter.
4. Write an equation which relates the values of the two variables.
5. If the line is extended, will it pass through the origin? Why?
6. Using the information obtained, predict the circumference of a circle whose diameter is 25 cm.

1-4 • MEASUREMENT AND THE CONCEPT OF DENSITY

Precise measurement is very important in laboratory work. If you are unfamiliar with the available balances, practice weighing objects before trying to find the density of the objects provided for this experiment. Determine the degree of uncertainty for the balance so that you can use significant figures properly in recording weight.



An assortment of solid objects that could be used in this experiment.

You will be given three solids. Two will have regular dimensions; one of these will float in water. The third will have irregular dimensions. As part of this experiment you are to decide how you can find the volume of each object. Make a data table for recording your observations.

PURPOSE:

1. To practice making precise measurements.
2. To practice using significant figures properly.
3. To provide practical experience in finding the density of objects in various ways.

PROCEDURE:

1. Prepare your data table.
2. Obtain the three objects described above. Predict which object has the greatest density and which has the lowest. Record your predictions.

3. Find the mass of each object. Record it, using the proper number of significant figures.
4. Find the volume of each object by an appropriate method. Record it using the proper number of significant figures.

ANALYSIS OF DATA:

1. Calculate the density of each object by applying the definition for density.
2. Find out the identity of the objects from the teacher. Look up the accepted value for the density of each object in a *Handbook of Chemistry and Physics*.
3. Find the experimental error and the percentage error for each density determination.
4. Compare your predictions with your results. Propose a hypothesis to account for any differences.
5. Which method for determining volume is the most precise? Why?
6. If several students had samples of the same material, but in widely different volumes, how could you check the accuracy of their results?
7. Refer to the slope of the graph you plotted for Experiment 1-3, Part I. What is the density of water at room temperature?

1-5 • HEAT ENERGY AND RANDOMNESS

You have observed sugar dissolving in water at room temperature. As the dissolving took place, the orderly arrangement of the sugar cube was destroyed. The particles of sugar became randomly distributed in the water and were no longer visible.

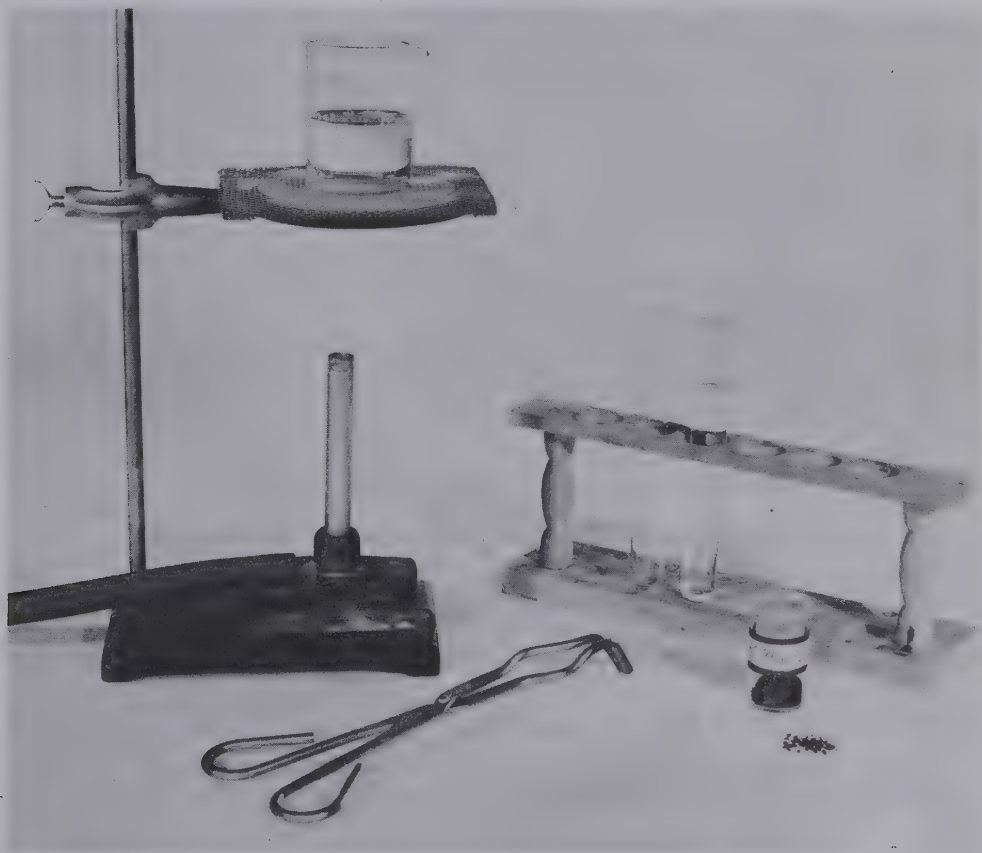
If we increase the water temperature, would the rate of dissolving change? In this experiment, a colored solid lets you follow the process of dissolving more easily. The solid is potassium permanganate, KMnO_4 . It exists in rhombic crystals as a solid.

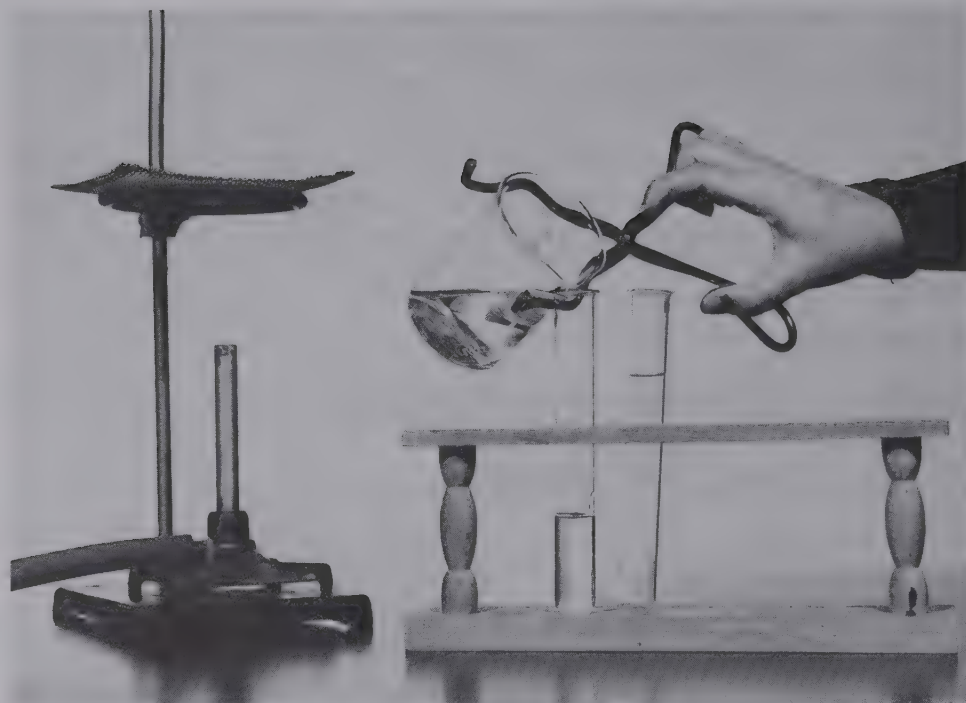
PURPOSE:

1. To predict the effect of heat energy on the rate at which a crystalline solid becomes randomly distributed in water.
2. To test your prediction by experiment.

PROCEDURE:

1. Prepare a data table.
2. Select two pinhead-sized pieces of potassium permanganate. They should be as nearly the same size as possible.
3. Label two 15 mm \times 150 mm test tubes: No. 1, room temperature water; No. 2, hot water.
4. Fill tube No. 1 to within 2 cm of the top with room temperature water.





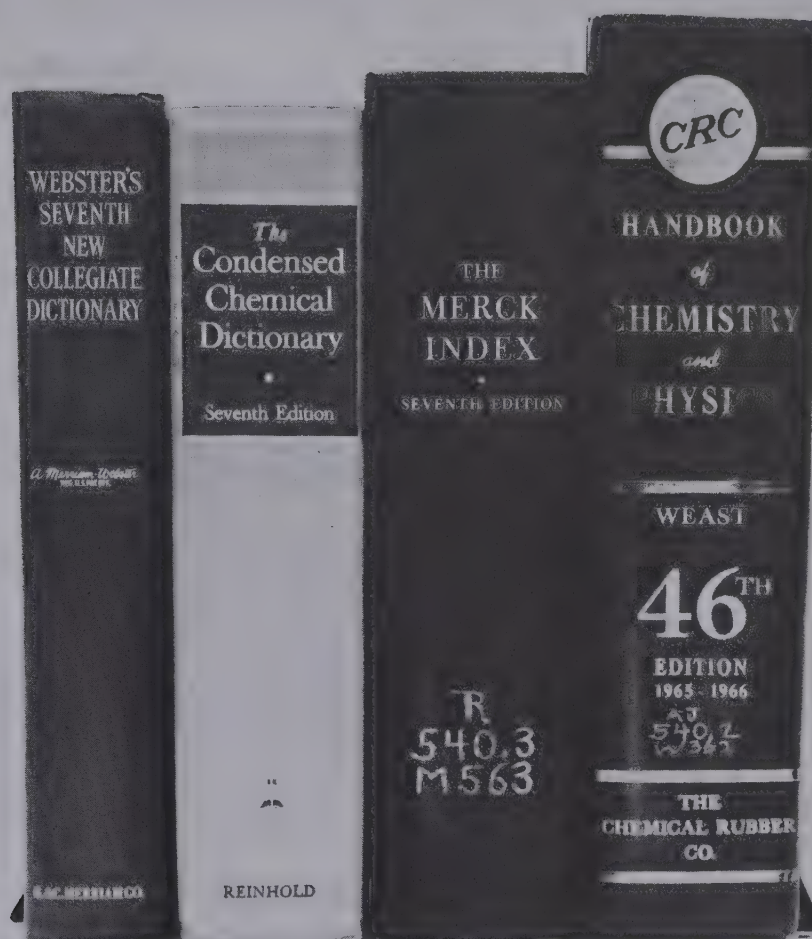
5. Place about 50 cc of water in a 100 cc beaker and heat it to boiling. While the water is heating, note and record the temperature of the water in the No. 1 tube.
6. When the water begins to boil, use beaker tongs to pick up the beaker and pour the water into the No. 2 tube until it is at the same level as the water in the No. 1 tube. (CAUTION: Handle boiling water carefully.)
7. Note and record the temperature of the water in tube No. 2.
8. Record your predictions in the laboratory record book: a. Will the solid dissolve more rapidly in hot water or in water at room temperature? b. In which case will the color of the dissolved solid be distributed more quickly and uniformly?
9. Drop one potassium permanganate crystal into each of the test tubes at the same time. Record observations for 15 minutes. Do *not* disturb the tubes during this time.

QUESTIONS:

1. In which test tube was more heat energy available to the solid?
2. In which test tube was the color distributed more rapidly?
3. What effect did heat have on the rate of dissolving?
4. On the basis of these observations, does heat energy favor order or randomness?
5. What do you think would happen if a crystal of potassium permanganate were placed in water at 6°C? Would it dissolve rapidly? Would the color be uniformly distributed? Give reasons for your answers. Check your answers by further experimentation.

EXERCISE 1-1 • LEARNING TO USE A CHEMISTRY HANDBOOK

One of the most useful references for the chemistry student is a chemistry handbook. Such books, differing slightly in format and content, are published by several companies. This exercise is designed to introduce you to one of the most popular of these books, the *Handbook of Chemistry and Physics*. However, if this book is unavailable in your school, you can do the exercise with another handbook.



Why is a dictionary placed with the three scientific reference books?

Handbooks are encyclopedic and technical. This means that they contain much information you will never use. The information you do use, however, is important.

PURPOSE:

To learn how to find needed information in the *Handbook of Chemistry and Physics*.

PROCEDURE:

1. Secure a copy of the *Handbook of Chemistry and Physics*. Note the full name of the publisher and the date of publication.
2. Read the table of contents to get an idea of the material contained in the handbook.
3. Look at the index. Learn how to locate information you may need.
4. The section "Physical Constants of Inorganic Compounds" can be especially helpful as you study chemistry. Turn to this section and list the information given about each substance.
5. Using *only* the handbook, answer each question below. List the page on which the answer is found and the title of the table or section used.
 - a. What is the atomic radius of aluminum?
 - b. What is the antidote for carbon monoxide poisoning?
 - c. Who discovered nickel?
 - d. How many grams of potassium permanganate will dissolve in 100 cc of water at 20°C?
 - e. What is the chemical formula for lead ferrocyanide?
 - f. What is the specific gravity (density) of sunflower seed oil?
 - g. What is the boiling point of glycerin?
 - h. What are the chemical hazards of carbon tetrachloride?
 - i. What is the cube root of 705?
 - j. How do you prepare a solution of congo red, a laboratory reagent?
 - k. What is the range of wavelengths of cosmic rays?
 - l. What is the definition of cryohydrate?
 - m. How many kilometers is Saturn from the sun?
 - n. How do you convert millimicrons to Angstrom units?

2-1 • CLASSIFICATION OF MATTER

In this experiment you will group samples of matter in several ways, including by appearance (homogeneous matter and heterogeneous matter) and by state (solid, liquid, and gas). Precise grouping may require testing which you cannot do at this time. Therefore, if you cannot decide upon the classification of a sample, place it in the group labeled "Unable to determine."

PURPOSE:

To gain experience in classifying matter according to different systems.

TECHNIQUES:

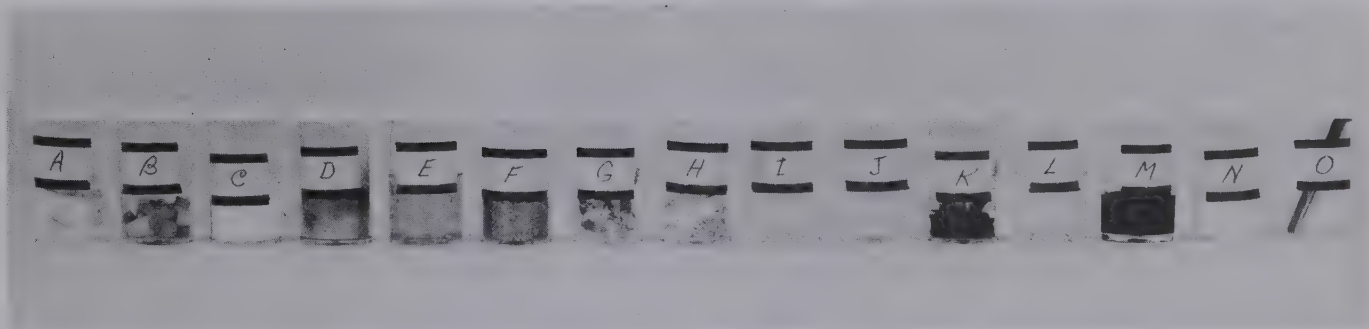
To test liquids with litmus paper and hydrion paper:

- Put a small piece of test paper on a glass slide.
- Dip a stirring rod into the solution to be tested and transfer one drop to the test paper.
- Observe for a change in color, if any.

PROCEDURE:

Part I:

- Examine the samples carefully. Separate them into two groups: (1) homogeneous matter and (2) heterogeneous matter.



- Separate the homogenous samples further into (1) solids and (2) solutions.
NOTE: To find out whether a liquid is a solution, place two or three drops of it on a clean glass slide. Hold the slide above a flame until the liquid evaporates. If a residue remains, the original liquid contained a dissolved solid. It was a homogeneous mixture (solution).
- Examine the solids again. Separate them into three groups: (1) elements, (2) compounds, and (3) unable to determine.
- Examine the solids once more and separate them into three groups: (1) metals, (2) nonmetals, and (3) unable to determine.
- Test each liquid with neutral litmus paper. Separate them into three groups: (1) acids, (2) bases, and (3) unable to determine. NOTE: If the litmus turns blue, the liquid is a base; if red, the liquid is an acid. If it does not change color, list the liquid in the group "Unable to determine."

Part II:

1. Heterogeneous matter is usually classed as a mixture. Examine the heterogeneous matter once again and select one mixture.
 - a. Think of a way to separate this mixture into its parts.
 - b. Have your idea approved by your teacher.
 - c. Go ahead and separate the mixture.
2. Select one solid from the "Unable to determine" group in Part I, step 3.
 - a. Propose a hypothesis about whether it might be an element or a compound.
 - b. Devise a way to test your hypothesis.
 - c. Have your method approved by your teacher.
 - d. Test your hypothesis by the method you have had approved.

QUESTIONS:

1. What other properties might be used to group these samples? Why do you think these would be useful methods of grouping?
2. List at least two reasons why it is hard to classify matter properly by inspection. Use evidence from this experiment.
3. Give one example of each of the following. (Do *not* use any of the samples you worked with.)

a. homogeneous mixture	c. solid	e. gas
b. heterogeneous matter	d. liquid	

2-2. PHYSICAL AND CHEMICAL PROPERTIES OF MATTER

Each kind of matter has certain characteristics that distinguish it from all other kinds of matter. These characteristics are known as properties. The chemist finds it useful to classify properties into two types: (1) physical and (2) chemical.

Physical properties are those which can be investigated without changing the chemical composition of the substance. *Chemical properties* are those which describe how the substance reacts with other substances to form a new substance with a different chemical composition.

PURPOSE:

To study the properties of aluminum foil so that the difference between physical and chemical properties will be more meaningful.

PROCEDURE:

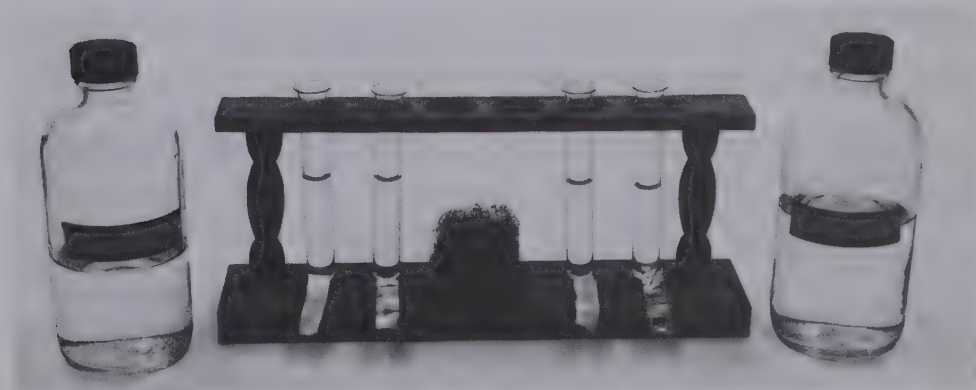
Part I: Physical Properties

1. Obtain a piece of aluminum foil. Record your observations on the following physical properties in your notebook:
 - a. length (use significant figures)
 - b. width (use significant figures)
 - c. mass (use significant figures)
 - d. color
 - e. odor
 - f. ability to dissolve in water
 - g. ability to conduct heat

Part II: Chemical Properties

Some of the evidences of chemical change are: (a) appearance of a gas; (b) disappearance of a substance; (c) appearance of a substance; (d) absorption of energy; (e) production of light, heat, or sound; (f) changes in color; (g) explosion; (h) presence of a new odor. Predict which of these you think might be observed while you perform this experiment.

1. Obtain a 6 cm \times 4 cm piece of aluminum foil. Cut it into two equal pieces. Rub one piece vigorously on both sides with steel wool. Cut it into 0.5 cm \times 4 cm strips.
2. Also cut the unrubbed piece of foil into 0.5 cm \times 4 cm strips.



3. Prepare four 13 mm \times 100 mm test tubes as follows:
 - a. Pour 5 cc of 6M HCl into two of the test tubes. Label.
 - b. Pour 5 cc of 6M NaOH into the other two test tubes. Label.
4. Place one rubbed strip of foil in one test tube of NaOH, and one unrubbed strip into the second test tube of NaOH. Record your observations.
5. Light a gas burner. Have wood splints and a clean, dry 13 mm \times 100 mm test tube available. (Review the tests for common gases in Experiment 1-2.)
6. Place one rubbed strip of foil in one test tube of HCl. Catch the gas evolved in the empty test tube by placing it over the reacting tube. When all action has stopped, light the splint and bring it to the mouth of the tube of gas. Record your observations.
7. Place an unrubbed strip of foil in the second tube of HCl. Record your observations.
8. Place 5 cc of distilled water in an evaporating dish. Then add the contents of one of the tubes containing HCl to the dish. Evaporate the liquid by placing the dish on an asbestos mat above the burner. Do not inhale the fumes. Note the color of the residue.
9. When the dish has cooled, add 3 cc of distilled water. Stir to dissolve the solid. Pour the solution into a clean test tube and add 3 cc of ammonia water. Record your observations.

QUESTIONS:

1. Using information gathered in this experiment, explain why such properties as color, mass, and odor are classed as physical properties. Why are such properties as the ability to react with acids and bases chemical properties?
2. List at least three evidences of chemical change that you observed in this experiment.
3. Propose a hypothesis to explain any difference in the reactions observed when the rubbed foil was placed in the solutions as compared with the way the unrubbed foil reacted.
4. What gas is produced when aluminum reacts with HCl? Give evidence for your conclusion.

QUESTIONS TO PONDER:

1. Would the concentration of the acid and base affect the rate of the reactions with aluminum foil?
2. Where does the gas come from?
3. How can the temperature change be explained?
4. What is the name of the residue obtained after the liquid is evaporated?
5. What is the name of the new substance produced when ammonia water is added to the residue?

2-3· MELTING AND FREEZING OF SOLIDS



Many substances which exist in the solid state at room temperature can be changed into the liquid state by the addition of heat energy. When heat is removed, the liquid again solidifies. A pure substance has a definite melting point. If the melting is gradual over a range of temperatures, the substance is probably a mixture.

In this experiment you will melt and freeze a solid and plot a time-temperature graph. It is best to work as partners. One partner can be the observer and the other the recorder for Part I of the experiment and vice versa for Part II.

Plan to plot both sets of data for the same solid on the same graph. Make two copies of the graph, one for each of you.

After completing the experiment using paradichlorobenzene, you will be given an unknown solid and asked to determine whether it is a pure substance or a mixture.

PURPOSE:

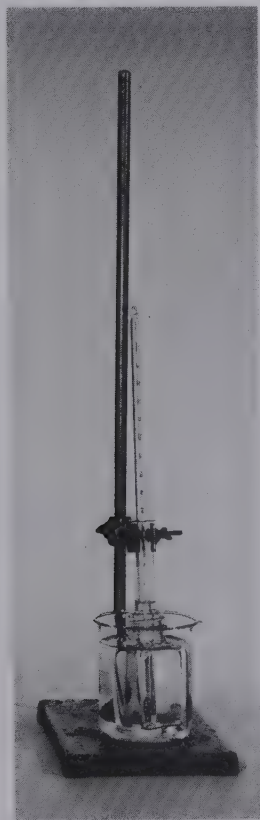
To study the energy changes as solids melt and freeze and to explain the observed results.

PROCEDURE:

Part I: Change of State from Liquid to Solid

Note: Make data charts in your record book before the period in which you plan to do the experiment. Allow space for recording the temperature every 30 seconds for 25 minutes. The data table should include a section for both cooling and melting.

1. Obtain a test tube of paradichlorobenzene from the instructor.
2. Attach a buret clamp to the neck of the test tube and attach it to a ring-stand.
3. Place a 60 cc gas collecting bottle in a 250 cc beaker. Add enough water to the beaker to come to the shoulder of the bottle.
4. Heat the test tube of solid by moving a *low* burner flame back and forth under it until the solid is completely melted. (CAUTION: Do not overheat the tube; heat it only enough to melt its contents.)
5. Place the beaker and gas collecting bottle under the test tube. Then lower the test tube until it rests on the bottom of the bottle. Insert the thermometer and start recording the temperature of the liquid as it cools.
6. Continue to record the temperature every 30 seconds until solidification is complete. Record when solidification begins, when it is complete, etc. Hold the thermometer just off the bottom of the test tube and against one side as the solid forms.



Part II: Change of State from Solid to Liquid

1. Raise the test tube of solid to the top of the ringstand and clamp it away from you.
2. Adjust a ring on your ringstand so that an asbestos mat placed on the ring is 3 cm from the top of the burner.
3. Light a burner. Adjust the flame so that the tip of it touches the asbestos mat. Place a 250 cc beaker about $\frac{2}{3}$ full of water on the mat and heat the water to about 95°C.
4. Record the temperature of the solid paradichlorobenzene. Then lower the test tube of solid into the gas collecting bottle which has been placed in the beaker of hot water. Keep the water temperature between 85°C and 90°C by heating if necessary. Record the temperature of the solid every 30 seconds until melting is complete. Note when it begins to melt and when melting is complete.
5. Remove the thermometer and clean it. Return all materials to the place designated by your teacher.

ANALYSIS OF DATA:

1. Plot a time-temperature graph of both sets of data, cooling and melting. Again, plot the independent variable on the x-axis; the dependent, on the y-axis. Plot both sets of data on the same graph. Use a tiny circle for each cooling temperature and a tiny x for each melting temperature. Join the points for the cooling curve with a black line; those for the heating curve with a red line.
2. Examine the graph and determine the melting temperature for the solid.
3. What is the freezing temperature of the solid?
4. Does a flat region, or plateau, occur in the graph? If so, propose a hypothesis to explain it.
5. Apply your knowledge of the opposing tendencies of minimum energy and maximum randomness to the melting and freezing of solids. Consider the following as you do so: (a) Can the particles of paradichlorobenzene be arranged in more ways when in the solid state or in the liquid state? (b) What did you do to the solid to melt it? (c) Which contains more energy, the solid or the liquid? (d) When the particles were most random, were they in a state of higher or lower energy?



2-4. ELEMENTS, MIXTURES, AND COMPOUNDS

Elements, mixtures, and compounds are not easily distinguished by inspection. In this experiment you will study the properties of two elements, iron and sulfur. Then you will compare the properties of a mixture of these elements with those of the separate elements. Lastly, you will prepare a compound composed of these same two elements and study its properties.

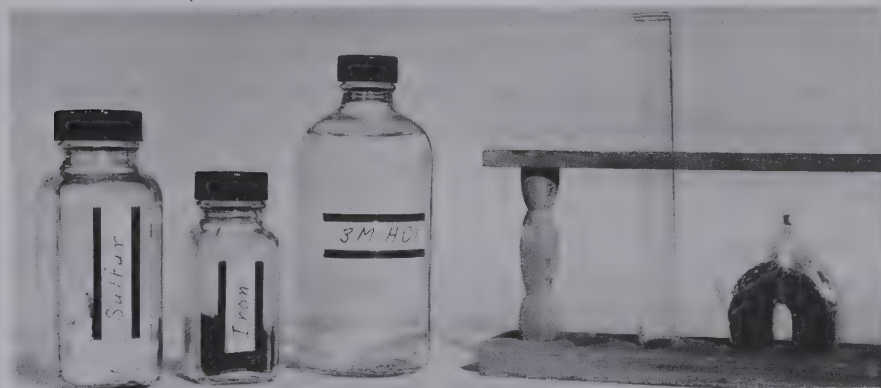
PURPOSE:

To observe the properties of two elements, a mixture, and a compound.

PROCEDURE:

Part I: Properties of Iron and Sulfur

1. Obtain a small piece of roll sulfur. Examine it carefully, recording as many physical properties as possible: color, odor, solubility in water, response to a magnet, etc. (Review the procedure for testing a water sample of a dissolved solid.)



2. Crush the lump of sulfur and place a bit of it in a test tube. Add 2 cc of HCl. Is there evidence of reaction?
3. Place about a half teaspoon of iron filings on a piece of paper. Record the physical properties.
4. Put some iron filings in a test tube. Add 2 cc of HCl. Is there evidence of reaction? If so, what products may be formed? Conduct whatever tests you can that may provide clues about what is happening. Record all observations.

Part II: Properties of a Mixture of Iron and Sulfur

1. Mix about a half teaspoon of crushed sulfur and a half teaspoon of iron filings on a piece of paper. Compare the physical properties of the mixture with those of the separate elements.
2. Place a pinch of the mixture in a test tube. Add 2 cc of HCl. Compare the results obtained with the results when HCl was added to the separate elements.
3. Filter the mixture and test the filtrate for dissolved solids.

Part III: Making the Compound Iron Sulfide

1. Weigh out 7 grams of iron filings and 4 grams of sulfur. Mix these thoroughly on a piece of paper.
2. Pour the mixture into a 15 mm \times 150 mm test tube.
3. Fill a 250 cc beaker $\frac{3}{4}$ full of tap water. Set it aside.
4. Heat the test tube strongly until the contents begin to glow. Observe the tube closely and when "a light comes on" in the lower part of the tube, remove it from the flame and watch it until the "light goes out." Then plunge the lower part of the test tube into the beaker of cold water. (Note: If the test tube did not break when it was plunged into cold water, wrap it in a towel to prevent flying glass, and tap it with a hammer to break it.)
5. If possible, weigh the compound for comparison with the 11 grams of mixture with which you began this part of the experiment.
6. Place a piece of the compound on a piece of paper and record as many physical properties as you can.
7. Place a small piece of the compound in a test tube. Add 2 cc of HCl. Observe the results. Compare these results with the action of each of the elements with HCl. Is there evidence of a chemical reaction? Record all observations. Note especially any odor for comparison with results in the early part of this experiment.

QUESTIONS:

1. Prepare a chart to summarize the major differences between mixtures and compounds which you observed in this experiment.
2. In which test or tests was a gas produced? If you obtained a gas in more than one test, try to determine whether it was the same gas. Give evidence for your conclusions.
3. Would other weights of iron and sulfur produce iron sulfide? (Testing answers to this question would be a good extension of this experiment if properly planned.)
4. What was the source of light energy which you observed as the elements reacted?
5. Why was it necessary to heat the mixture to start the chemical reaction?
6. Assuming that 11 grams of iron sulfide was produced by combining 7 grams of iron and 4 grams of sulfur, calculate: (a) the percentage of iron in the compound; (b) the percentage of sulfur in the compound; and (c) the simplest formula for iron sulfide.

Hint for part c:

Element	Percentage in iron sulfide	\div	atomic weight	=	relative number of atoms (round off)
iron	_____	\div	56	=	_____
sulfur	_____	\div	32	=	_____

Calculated simplest formula for iron sulfide = _____.

2-5 • FORCES BETWEEN LIKE AND UNLIKE MAGNETIC POLES

You have probably played with magnets enough to know that like poles repel and unlike poles attract. You may also know that the distance between the two poles affects the forces of both attraction and repulsion. The forces that act between poles can be studied more precisely by the technique of this experiment.

As you perform this experiment, consider very carefully how the potential energy changes in each situation that you set up.

The data you obtain and the graph you plot will be of great benefit to you when you begin your study of chemical bonding.

PURPOSE:

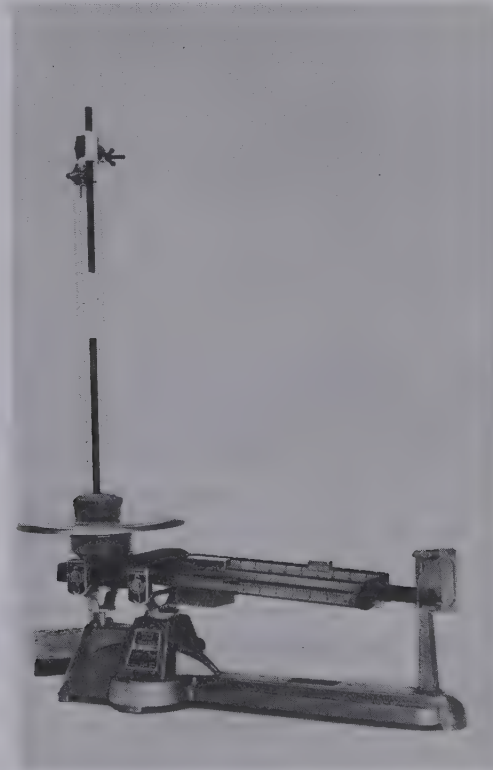
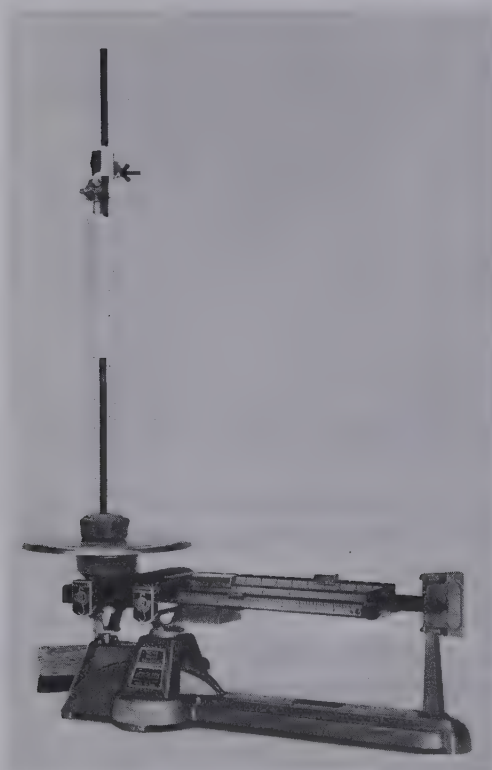
To obtain quantitative data about the relationship between force and distance between like and unlike magnetic poles.

TECHNIQUES:

Be sure the platform of the balance does *not* respond to the magnet. Set up the apparatus as shown in the illustration.

PROCEDURE:

1. Study the steps in this procedure and prepare a data table before you start the experiment. The readings will then be recorded in a meaningful manner.
2. Level the platform balance.



3. Obtain a pair of magnets. Stand one magnet on end on the platform as shown in the illustration. Secure it with modeling clay. Record the weight of this magnet and the clay.
4. Place a metric rule on a ringstand, using tape to secure it. Place it so that the scale can be used to observe the relative positions of the two magnets.
5. Clamp a second magnet to the ringstand so that *like* poles of the two magnets are facing. Move the clamp and magnet up or down until the two magnets are about 20 cm apart.
6. Lower the magnet held on the ringstand until it causes the balance reading to change. Record the distance between the two magnets in cm and the apparent weight in grams.
7. Continue to move the top magnet closer to the magnet on the balance, recording the distance and apparent weight each time until eight to ten readings have been taken.
8. Turn the magnet over in the clamp so that *unlike* poles are facing.
9. Again, begin with the movable magnet about 20 cm from the magnet on the balance. Lower it, and then record the distance and apparent weight until eight to ten readings have been taken.
10. Take the apparatus apart and put it away. Be sure to remove all the clay from the magnet.

ANALYSIS OF DATA:

1. Calculate the magnetic force in grams for each reading by subtracting the actual weight of the magnet and clay on the balance from the recorded, apparent weight of each reading. When the actual weight is less than the apparent weight, the force has a positive value. When the actual weight is greater than the apparent weight, the force has a negative value.
2. Draw a line across the middle of a sheet of graph paper. Label this line "zero magnetic force." Consider distance to be the independent variable and force to be the dependent variable. Plot positive magnetic force values above the zero line and negative magnetic force values below the zero line.
3. Join the points in each section of the graph with a smooth line.

QUESTIONS:

1. Which combination of poles gives positive magnetic force values? Which gives negative magnetic force values?
2. When your force values were negative, the upper magnet was pulling up on the lower magnet. Would you call such magnetic forces repulsive or attractive?
3. What relationship can you establish between the strength of a magnetic force and the distance between two facing magnetic poles?
4. Propose how the potential energy is changed as the positions of two magnetic poles are changed.

3-1 • LEARNING TO RECOGNIZE DYNAMIC EQUILIBRIUM

One of the most important concepts in chemistry is that of equilibrium. Equilibrium is a dynamic state in which two opposing processes occur simultaneously at equal but opposite rates. For equilibrium to be established, no matter or energy must enter or leave the system. Therefore, the system must be closed and not be heated or cooled. Equilibrium exists in a closed system when you can no longer see changes in the properties of the system.

In this experiment you will set up the following systems: (1) liquid-vapor, (2) solid-liquid, (3) gas-liquid, and (4) solid-gas. You will set up an open and a closed system in each case and compare the differences over time.

In the gas-liquid system, a chemical indicator (phenolphthalein) will be used. This indicator turns red when ammonia gas dissolves in water.

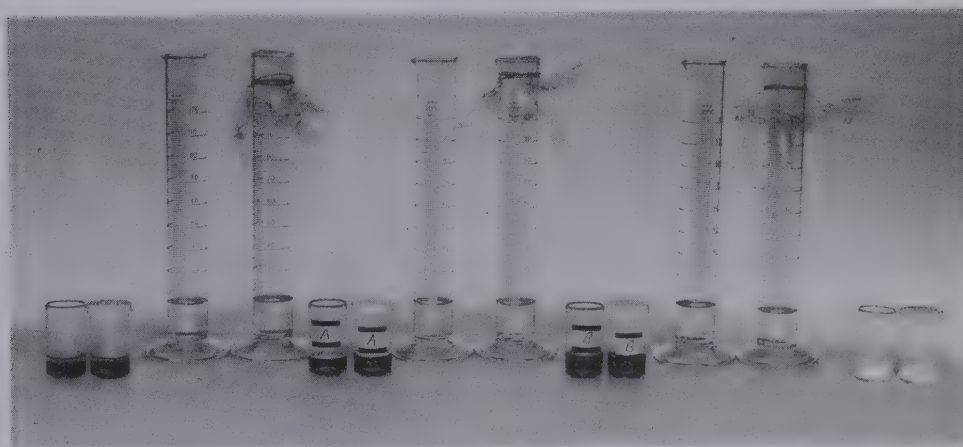
PURPOSE:

To observe open and closed systems to learn how to recognize dynamic equilibrium.

PROCEDURE:

Part I: Liquid-Vapor Systems

1. Obtain six 100 cc graduated cylinders. Label two "water," two "alcohol," and two "acetone."
2. Pour 5 cc of water into two cylinders. Seal one with a piece of plastic and a rubber band. Leave the other uncovered.



Which of the systems shown are equilibrium systems and which are not?

3. Repeat step 2 using alcohol and acetone.
4. Observe and record the levels of the liquids as often as possible for the next 24 hours.

Note: the vapor pressure of these liquids in torr at 20°C is:

water	17.5
alcohol	43.9
acetone	184.8

Part II: Solid-Liquid Systems

1. Obtain two 2 cm \times 4 cm flat-bottom vials and one 250 cc beaker.
2. Place a large crystal of copper sulfate in each vial. Add enough water to cover the crystals.
3. Cover one vial with the 250 cc beaker. Leave the second uncovered.
4. Record observations for about 15 minutes. You may go on to Part III, but observe these vials at intervals.

Part III: Gas-Liquid Systems

1. Obtain four 2 cm \times 4 cm flat-bottom vials, and one 250 cc beaker.
2. Label two vials "A" and two "B."
3. Pour 3 cc of ammonia water into each "A." Then add 1 drop of phenolphthalein and stir.
4. Pour 3 cc tap water into each "B." Add 1 drop of phenolphthalein and stir.
5. Place one "A" and one "B" vial close together. Cover them with the 250 cc beaker.
6. Place the other "A" and "B" vials close together and leave them uncovered.
7. Record observations on the two sets of vials for about 30 minutes.

Part IV: Solid-Gas Systems

1. Obtain two 2 cm \times 4 cm flat-bottom vials and one cap. Place about 1 cc of solid ammonium carbonate in each vial.
2. Seal one vial; leave the other open.
3. Cautiously smell the open vial.
4. Place both vials in a quiet place. Observe them every 24 hours for several days.

QUESTIONS:

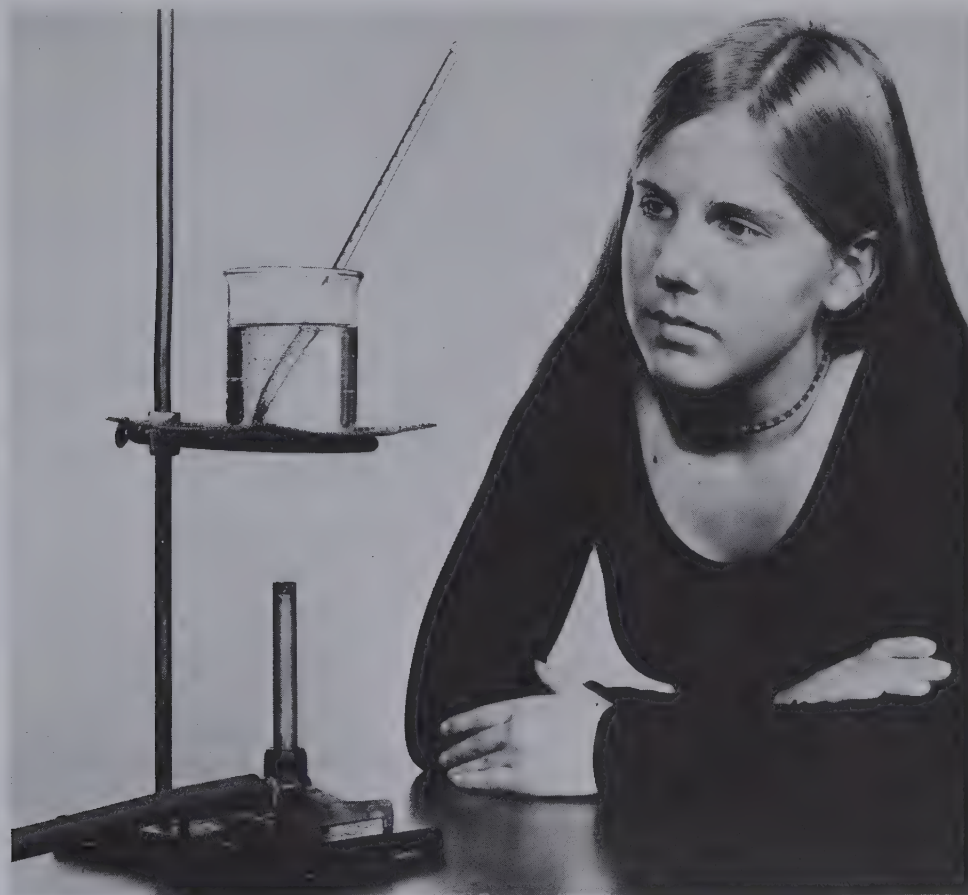
1. Why must a system be closed for equilibrium to be established?
2. How can you recognize an equilibrium system?
3. Dynamic equilibrium means that two processes are occurring at the same time and at the same rates. What two processes were occurring at dynamic equilibrium in each part of this experiment?
4. Propose a hypothesis to account for the volume changes observed in Part I.
5. If sealed tubes like those in Part I were set in the hot sunshine and in a cold refrigerator at the same time, which system would reach equilibrium first? Would you expect other differences in the two systems?
6. What would raising the temperature do to the systems in Parts II, III, and IV? Would the time required for equilibrium to be established be affected? Would the appearance of the systems at equilibrium be changed?
7. Refer to your observations from Experiment 1-5. Did either test tube in that experiment represent an equilibrium system? Give reasons for your answer.

3-2. THE PROCESS OF BOILING

You have seen water boil many times, but how much do you really know about the process of boiling? Like many events with which you are familiar, details go unnoticed. In this experiment you will observe carefully the changes that occur as water is heated from room temperature to the boiling point.

PURPOSE:

To observe changes that occur in water as it is heated to its boiling temperature.



PROCEDURE:

1. Fill a 250 cc beaker $\frac{3}{4}$ full of distilled water.
2. Adjust an asbestos mat on your ringstand so that it is about 6 cm above the top of your burner. Light the burner. Adjust the flame so that the tip just touches the asbestos mat.
3. Record the temperature of the water and set the beaker of water on the asbestos mat. Record the temperature every 2 minutes as the water is heated and for 6 minutes after it starts boiling.
4. As the water heats, observe the currents in the beaker and the size of bubbles as they appear. Follow the paths of some bubbles to see what happens to them as they rise.
5. Record the barometric pressure.

ANALYSIS OF DATA:

1. After you decide whether time or temperature is the independent variable, graph your data. Draw a smooth line which connects most of the points.
2. What is the slope of the line? Explain the meaning of the slope in terms of the two variables in this experiment.
3. Is there a plateau on your graph? If so, propose a hypothesis to account for it.

QUESTIONS:

1. Were the first bubbles observed large or small? What do you think they are?
2. As the water neared the boiling point, were the bubbles larger or smaller than the first ones observed?
3. Why did some bubbles disappear before reaching the surface of the water?
4. Did the vapor pressure of the water change as the temperature increased? Give evidence to support your answer.
5. If the pressure on the surface of the water were increased, how do you think this would affect the boiling temperature of the water? What household appliance makes use of this principle?
6. Will water boil at the same temperature at the base of Pike's Peak and at the top of this mountain? Support your answer.
7. Will the boiling temperature of water change if solids are dissolved in it?
8. Formulate a definition of *boiling* based on your observations in this experiment.

SUGGESTIONS FOR AN ADDITIONAL INVESTIGATION:

Propose a hypothesis to explain why the bubbles rise from specific points in the beaker. Devise an experiment to test your hypothesis. Pour a carbonated beverage into a glass and observe the origin of the bubbles.

3-3. THE CHEMICAL PROPERTIES OF WATER

This experiment investigates reactions between water and various substances. Much of the information gained will be used later in the study of acids and bases. As you observe the reactions, think about the type of chemical reaction taking place. Write a word equation for each reaction.

PURPOSE:

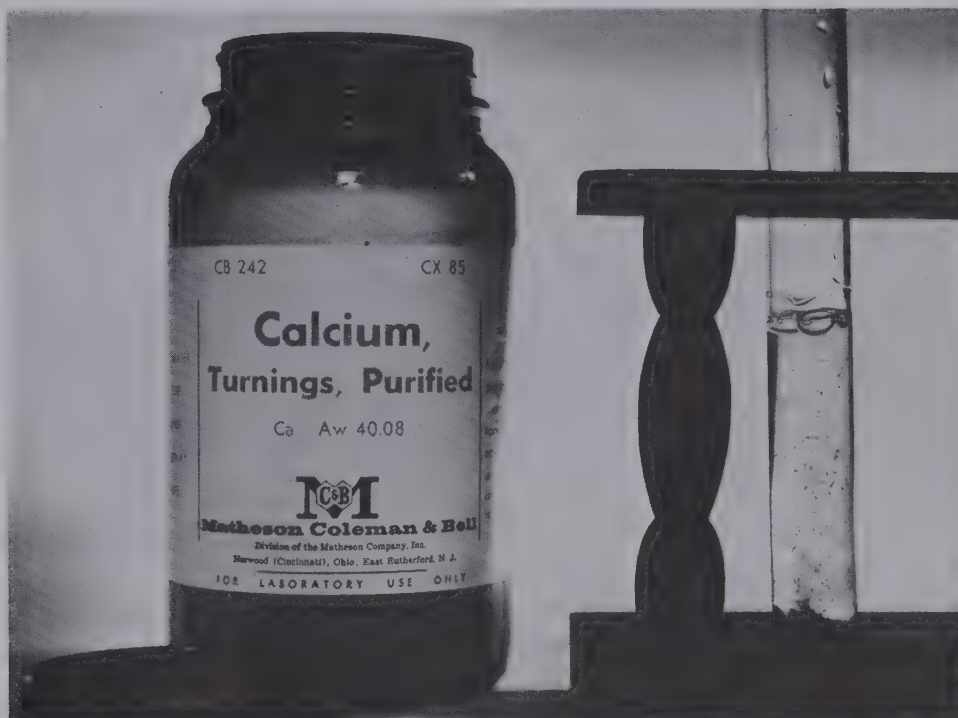
To observe the chemical properties of water and to represent some reactions with word equations.

PROCEDURE:

Part I: Reaction between Water and an Active Metal



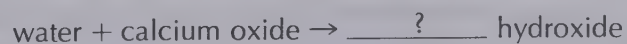
1. Pour about 2 cc water into a large test tube. Drop a small piece of calcium metal into the water. Record observations.



A piece of calcium metal has been placed in water in the test tube.

2. Cap the test tube with your thumb for a few seconds, then bring a burning splint to the mouth of the test tube. What gas is being produced?
3. Test the solution in the test tube with litmus paper. Record results.

Part II: Reaction between Water and a Metallic Oxide



1. Place a small pile of freshly opened calcium oxide on a watch glass. Slowly add about 2 cc of water. Is heat released?
2. Test the solution with litmus paper. Is it acidic or basic?

Part III: Reaction between Water and a Nonmetallic Oxide

water + sulfur dioxide \rightarrow hydrogen sulfite (also called sulfurous acid)



What reactions are occurring inside the gas-collecting bottle?

1. Add water to a gas collecting bottle to a level of about 1 cm. Obtain a small glass plate to cover the bottle.
2. Put a piece of roll sulfur about the size of a pea in the bowl of a deflagrating spoon. Ignite the sulfur by holding the bowl of the spoon in the burner flame. As soon as the sulfur begins to burn, quickly lower the bowl of the spoon into the gas collecting bottle. Do *not* let the bowl of the spoon touch the water. Cover the mouth of the bottle with the glass plate to prevent escape of any gas produced. When the sulfur has stopped burning, remove the spoon from the bottle. Recover the bottle with the glass plate. Shake it to dissolve as much of the gas as possible.
with the glass plate. Shake it to dissolve as much of the gas as possible.
3. Test the resulting solution with hydriion paper. Is it acidic or basic?

Part IV: Water Promotes Chemical Reactions

lead nitrate + potassium iodide \rightarrow ____?____ iodide + potassium ____?____

1. Mix together on a clean, dry watch glass a few crystals of lead nitrate, $\text{Pb}(\text{NO}_3)_2$, and potassium iodide, KI. Is there evidence of a change in the crystals as they are mixed?
2. Add a few drops of water to the mixture. Record your observations.

QUESTIONS:

1. Complete the word equations for each reaction.
2. On the basis of your observations list in general terms four chemical properties of water.
3. Calcium oxide is often called lime. Calcium hydroxide is called slaked lime. Slaked lime is sometimes used to line football fields. How might it be dangerous to perspiring players if the field were lined with lime? Check the dictionary for the meaning of *slake*.
4. Why would you expect unpainted iron to rust sooner than painted iron?
5. A girl used some baking powder to make biscuits. The can of baking powder had not been tightly sealed. Why didn't her biscuits rise properly?

3-4. ENERGY CHANGES IN SOLUTIONS

In Experiment 1-2 you watched a sugar cube dissolve in water. In this experiment you will take another look at the solution process, giving special attention to the energy changes that occur.

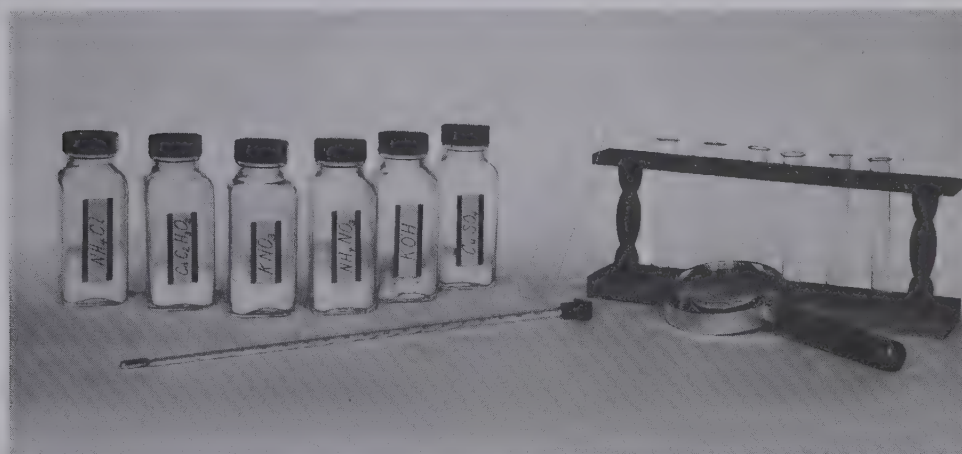
PURPOSE:

To observe the process of solution and to study the energy changes that occur.

PROCEDURE:

Note: Study each of the solids one at a time so that you can give your full attention to each. This means that you will repeat steps 2 through 8 for each solid you study.

1. Obtain about 1 gram of each of these solids:
 - a. ammonium chloride, NH_4Cl
 - b. calcium acetate, $\text{Ca}(\text{C}_2\text{H}_3\text{O}_2)_2$
 - c. potassium nitrate, KNO_3
 - d. ammonium nitrate, NH_4NO_3
 - e. potassium hydroxide, KOH
 - f. anhydrous copper sulfate, CuSO_4



What are the possible things that could happen when these substances are put in water?

2. Pour about 5 cc distilled water into a 100 mm \times 13 mm test tube.
3. Sprinkle just a few crystals of the solid (one pellet in the case of KOH) into the water. Observe the solution process very carefully. Particularly look for any evidence of chemical change.
4. Record the temperature of the water.
5. With the thermometer in the water, quickly add the remainder of the solid and record the maximum temperature change.
6. Rinse and dry the thermometer.
7. Place a drop of the solution on the end of a glass slide. Evaporate the water.
8. Observe the residue under a strong magnifier, preferably a binocular microscope. Place a very few, *not* previously dissolved particles of the same solid on the same slide and compare the samples. Does dissolving appear to change the structure of the solid?
9. Repeat the process for the remaining solids.

QUESTIONS:

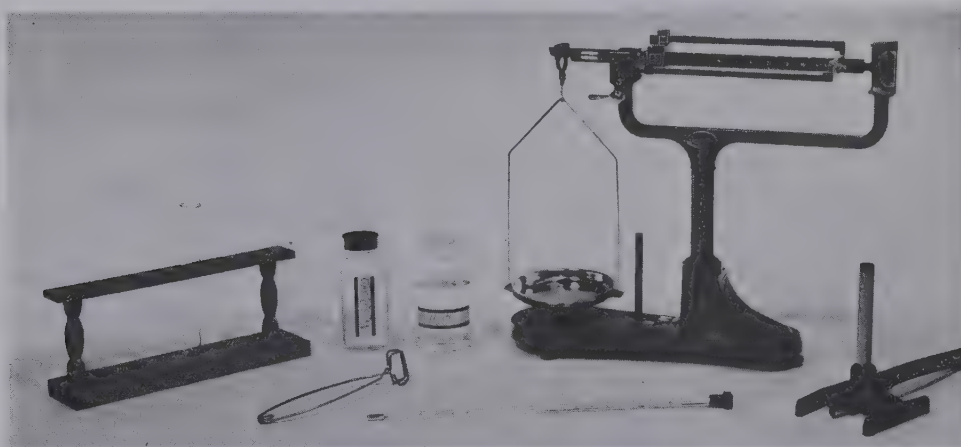
1. Which solids caused the temperature of the water to increase as they dissolved? Which caused the temperature to decrease?
2. Which solids do you think would be more soluble in hot water than in cold? Give reasons for your answer.
3. Propose a hypothesis to explain why some solids dissolve endothermically and why some dissolve exothermically. (Hint: Which solids apparently have stronger forces holding the particles together?)
4. Does the dissolving process seem to affect the structure of the solid? Support your answer.

3-5 • UNSATURATED, SATURATED, AND SUPERSATURATED SOLUTIONS

Aqueous solutions are unsaturated if more solute can be dissolved at the existing temperature and pressure. If the maximum amount of solute is in equilibrium with undissolved solid, the solution is saturated. If a hot, saturated solution is prepared and allowed to cool without losing any solute, the solution is supersaturated. Most solutes fall out of solution as the temperature is lowered, but a few substances remain in solution during the cooling process. One such solid is sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$. In this experiment you will prepare the three types of solutions using sodium thiosulfate. This solid is used extensively by photographers; they often call it "hypo."

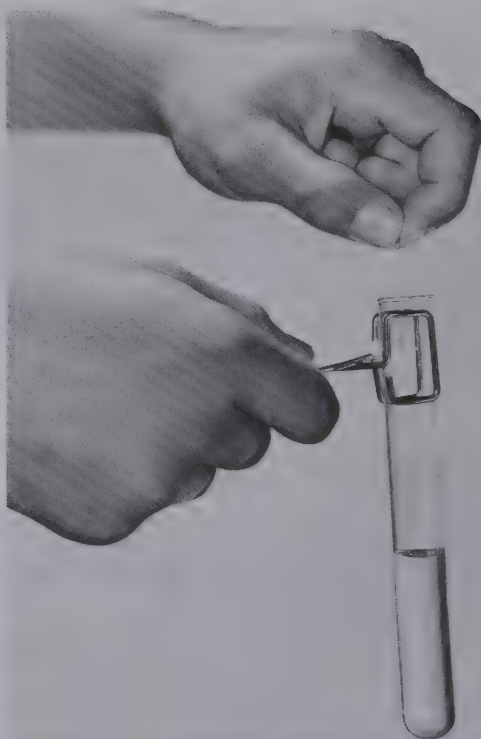
PURPOSE:

To study the properties of unsaturated, saturated, and supersaturated solutions.



PROCEDURE:

1. Weigh out 15 g of sodium thiosulfate on a piece of paper.
2. Pour a precisely measured 3 cc of distilled water into an 18 mm \times 150 mm test tube.
3. Add crystals of sodium thiosulfate a few at a time until no more will dissolve at room temperature. About 4 grams is required to produce a saturated solution. Shake the tube to speed the solution process. Note any change in temperature as solution occurs. Do not use a thermometer; it will remove some of the water, and the 3 cc is rather critical. To note the temperature change, touch the bottom of the test tube to your inner wrist from time to time.
4. As you heat the solution very gently (do *not* allow it to come to a boil), continue to add solid until the entire 15 grams is in solution. Remove the test tube from the flame and place it in a beaker of tap water. Let the test tube and contents stand undisturbed until they are at room temperature. This takes about 15 minutes.
5. Pick up the tube and examine the contents closely. The solution should be clear with no solid visible. Record the temperature of the solution.



6. Drop a tiny piece of solid sodium thiosulfate into the supersaturated solution and observe what happens. Record the temperature every 15 seconds until all action has stopped.
7. Dispose of the contents of the tube as directed by your instructor.

QUESTIONS:

1. How can such a large amount of solute be dissolved in only 3 cc of water? (Hint: Look at the formula for sodium thiosulfate.)
2. Did the dissolving process at room temperature raise or lower the temperature of the water? How can you account for this change in energy?
3. Did the process of crystallization absorb or release heat? How can you account for this change in energy?
4. Which is more random, the crystal or the dissolved solute? Give reasons for your answer.
5. Which solution represented the most stable state? The most unstable? Give evidence for your answers.
6. Which contains more potential energy, the crystal or the dissolved solute? Give a reason for your answer.

SUGGESTIONS FOR AN ADDITIONAL INVESTIGATION:

What sort of graph would be produced by the data, if you kept a time-temperature record of this experiment similar to the one you kept for Part II of Experiment 2-3? Do an experiment that tests your prediction.

4-1 • ELECTRICAL CONDUCTIVITY OF LIQUIDS

The water solutions of many compounds can conduct an electric current. The amount of current conducted depends upon the number of ions present in the solution. Solutions that are good conductors are called strong electrolytes. Those that conduct current poorly are called weak electrolytes. Solutions that conduct no current are nonelectrolytes.

Both the nature and the concentration of the dissolved compound affect the amount of current a solution will conduct. In this experiment we will investigate the conductivity of pure water, water solutions, and other liquids. This experiment is a foundation for the study of acids, bases, and salts.

PURPOSE:

To classify water solutions of compounds as strong electrolytes, weak electrolytes, and nonelectrolytes and to compare the conductivity of solutions with that of pure liquids.

PROCEDURE:



Set up equipment similar to that shown in the illustration. The meter shown contains an internal source of current. When you complete the circuit by placing the electrodes in a highly conductive solution, the needle swings to the right and off the scale as shown. For less conductive or nonconducting liquids it may move not at all or only slightly. When this happens, touch the electrodes together to make sure the circuit is complete in all other respects.

1. Prepare a data table using columns with these headings:

Liquid	Meter Reading	Strong Electrolyte	Weak Electrolyte	Nonelectrolyte
--------	---------------	--------------------	------------------	----------------

2. Observe the results and record the meter readings for each liquid and solution tested. The following list represents a good sample of liquids and solutions.

Distilled water

Tap water

Distilled water with a few tiny crystals of sodium chloride stirred into it

Sodium chloride (0.1M)	}	Salt solutions
Potassium nitrate (0.1M)		
Copper (II) sulfate (0.1M)		

Hydrochloric acid (0.1M)	}	Acids
Sulfuric acid (0.1M)		
Nitric acid (0.1M)		
Acetic acid (0.1M)		

Sodium hydroxide (0.1M)	}	Bases
Potassium hydroxide (0.1M)		
Ammonium hydroxide (0.1M)		
Calcium hydroxide (saturated)		

Alcohol	}	Organic compounds
Carbon tetrachloride		
Benzene		
Acetone		
Pure acetic acid		

Sugar + water	}	Organic solutions
Sugar + alcohol		
NaCl + alcohol		
HCl + benzene		
Pure acetic acid as water is added		

QUESTIONS:

1. What classes of compounds were strong electrolytes? Weak electrolytes? Nonelectrolytes?
2. What is done to control conditions in this experiment? Why is it important to do this?
3. Propose a hypothesis to explain the change in conductivity of pure acetic acid as water was added. How might you test your hypothesis?
4. Propose a hypothesis to explain why some substances are good conductors of electricity while others are poor conductors, or even nonconductors.

4-2. PREPARATION OF A SALT FROM A TYPICAL ACID AND BASE

Acids and bases have characteristic properties which identify them. The way they affect litmus paper and other chemical indicators, their taste, and their effect upon other substances are some of the properties that allow us to tell an acid from a base. Equivalent amounts of an acid and a base react to form equivalent amounts of salt and water. This process, known as neutralization, can be represented by the general word equation:



For example, table salt is sodium chloride. It is composed of positively charged sodium ions from a base (sodium hydroxide) and negatively charged chloride ions from an acid (hydrochloric acid). There are many other salts. Each of them is made up of positively charged and negatively charged ions which come from the corresponding ions in solutions of base and acid.

PURPOSE:

To study the properties of a typical acid and base and then to prepare a salt by combining equivalent amounts of each.



PROCEDURE:

Part I: Properties of Hydrochloric Acid, HCl

1. Pour 2 cc of 3M HCl into each of three test tubes (13 × 100 mm). Label them A, B, and C.
2. To test tube A add a small piece of neutral litmus paper. Record results.
3. To test tube B add 3–4 drops of phenolphthalein indicator solution. Results?
4. Drop a small piece of calcium carbonate, CaCO_3 , into test tube C. Record all observations.
5. What possible gases could have been formed in step 4? Conduct appropriate tests to identify the gas.

Part II: Properties of Sodium Hydroxide, NaOH

1. Pour 2 cc of 3M NaOH into each of three test tubes (13 × 100 mm). Label these A, B, and C.
2. Repeat the tests performed in Part I. Record observations.

Part III: Preparation of a Typical Salt, Sodium Chloride, NaCl, from Hydrochloric Acid and Sodium Hydroxide

1. Place 10 cc of 3M NaOH solution in a clean evaporating dish. Add 4 drops of phenolphthalein and stir.
2. Pour about 15 cc of 3M HCl into a clean graduated cylinder. Add the HCl, a few drops at a time, to the NaOH solution in the evaporating dish. Stir constantly. Continue to add HCl until the pink color disappears completely. CAUTION: When about 9 cc of the acid has been added, proceed by adding only one drop at a time.
3. When the solution is neutral, as indicated by the loss of the pink color, place the evaporating dish on an asbestos mat over a burner. Heat gently until the liquid has evaporated. If the pink color returns, add a few more drops of HCl solution. (Refer to the technique for evaporating a liquid described in the introduction to this laboratory manual.) Examine a few crystals of the residue with a hand lens.
4. Carefully note the taste of the residue.

Part IV: Preparation of NaCl from HCl and NaHCO₃

1. Put about 0.2 gram NaHCO₃ on a clean glass slide. NaHCO₃ is baking soda; a small pinch will do instead of weighing it on a balance.)
2. Add HCl drop by drop to the baking soda, as in Part III, until the reaction stops.
3. Evaporate to dryness by heating gently. Examine the residue with a hand lens.
4. You might wish to compare the taste of this residue with the taste of baking soda. Do so with care.

QUESTIONS:

1. List three properties of an acid observed in this experiment.
2. List three properties of a base observed in this experiment.
3. Write the word equation for the neutralization of hydrochloric acid with sodium hydroxide.
4. Suggest a method for preparing the following salts in the laboratory.
 - a. potassium chloride
 - b. sodium sulfate
 - c. potassium nitrate

2. If you do not have spot plates like those in the illustration use a grease pencil or a crayon to draw a grid on a piece of clear plastic that corresponds to the data table you have prepared. Make the squares at least $2\text{ cm} \times 2\text{ cm}$. Place the plastic sheet on top of a sheet of white paper.
3. Place 2–3 drops of acid or base of the indicated concentration in each square of the column for that concentration. Place 2–3 drops of distilled water in each square of the middle column.
4. Add the indicator to each square across the row for that indicator.
5. Record the colors produced in each square in your data table. Look carefully for any subtle color variations from square to square across a row.
6. Compare your results with the color plate on page 121 of the text.

QUESTIONS:

1. What effect does dilution have on the pH of an acid? on that of a base?
2. What is the expected pH of the following acid concentrations?
a. 0.00001M b. 0.01M c. 1.0M
3. What is the expected pH of the following base concentrations?
a. 0.1M b. 1.0M c. 0.0001M
4. If a very precise determination of the pH of a solution were to be made, which indicator would you choose?
5. Use reference books to find the composition of the indicators used in this experiment. If you do not already know where the common indicator, litmus, comes from, look this up also.
6. Suggest why all the possible dilutions of H_3O^+ and OH^- were not used for this experiment.

SUGGESTIONS FOR AN ADDITIONAL ACTIVITY:

Determine the pH of such common substances as: lemons, oranges, grapefruit, apples, bananas, grapes, vinegar, milk, carbonated soft drinks, blood, and other substances you can think of.

4-4 · HYDROLYSIS, THE COMMON ION, AND BUFFERS

Hydrolysis is a reaction that occurs when the ions of a dissolved salt react with water to form hydronium or hydroxide ions. If the ion from the salt which reacts is an acid, hydronium ions will be formed. If the ion from the salt which reacts is a base, it will take hydrogen ions, protons, from water molecules, thus producing hydroxide ions. If the salt is the salt of a weak base and a weak acid, then both of its ions will react with water. The solution will have a slight change in pH, usually, depending upon whether the parent acid or base is the weaker. So, in general, hydrolysis affects the pH of a solution.

Common ions also affect the pH through a buffering action. Take a solution already at equilibrium. If a salt is added which contains an ion like one of those already in dynamic equilibrium with its molecular form in solution, the equilibrium shifts so that some of the added common ion is converted to the molecular form. The resulting mixture of ions and molecules is a buffer solution. If either additional hydronium or hydroxide ions are added, the buffer solution can react with them so that the pH remains quite stable. A solution of a weak acid or a weak base with one of its salts added to the solution acts as a buffer solution.

In this experiment you will test the pH of various salt solutions to determine the effect of hydrolysis. You will also observe the effect on the pH when a common ion is added to a system in dynamic equilibrium.

PURPOSE:

To investigate the effect of certain salts upon the pH of water as a result of hydrolysis; to observe the effect of a common ion; and to see how a buffer controls pH.

PROCEDURE AND RESULTS:

Part I: Hydrolysis of Salts

1. Prepare a data table as shown:

1	2		3		4			5
Formula of Salt	Parent		Strength of		Predicted pH range			Tested pH
	Acid	Base	Acid	Base	Acidic	Neutral	Basic	

- Enter the formulas of the salts listed below in the first column:

sodium chloride, NaCl	potassium nitrate, KNO ₃
sodium carbonate, Na ₂ CO ₃	ammonium chloride, NH ₄ Cl
copper sulfate, CuSO ₄	sodium sulfate, Na ₂ SO ₄
sodium acetate, NaC ₂ H ₃ O ₂	sodium phosphate, Na ₃ PO ₄
aluminum sulfate, Al ₂ (SO ₄) ₃	



- Write the formulas for the parent base and parent acid in the second column.
- Indicate in the third column whether the parent base is strong or weak. Do the same thing for the parent acid.
- Enter your prediction about the pH of the salt solution in column 4.
- Test the pH of each salt solution with hydrion paper. Enter results in the last column. (Note: Be sure to test each solution by transferring a drop of the solution to a small piece of hydrion paper which has been placed on a clean glass slide or watch glass. Placing the hydrion paper in the solution contaminates the solution!)

Part II: Le Chatelier's Principle and the Common Ion

A. The Effect of the Common Ion in Equilibrium Systems in which MOLECULES Predominate

- Pour 10 cc of 1M acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$, into each of two test tubes, labeled #1 and #2. Find the pH with hydrion paper. Then add 3 drops of methyl orange to each test tube. Record the color of the solutions in a data table.
- Add 1 gram of solid sodium acetate, $\text{NaC}_2\text{H}_3\text{O}_2$, to test tube #1. Record the color and pH.
- Pour 10 cc of 1M ammonia water into each of two test tubes, labeled #3 and #4. Add 3 drops of phenolphthalein to each. Record the color and pH of the solutions in a data table.
- Add 1 gram of solid ammonium chloride, NH_4Cl , to test tube #3. Record the color and pH.

B. The Effect of the Common Ion in Equilibrium Systems in which IONS Predominate

- Pour 10 cc of 1M hydrochloric acid, HCl , into each of two test tubes, labeled #1 and #2. Add 3 drops of methyl orange to each test tube. Record the color and pH.
- Add 1 gram of solid potassium chloride, KCl , to test tube #1. Record the color and pH.

Part III: The Effect of Buffers on pH

1. Pour 5 cc of distilled water (pH of 7) into each of two test tubes, labeled #1 and #2.
2. Pour 5 cc of a buffer solution (pH of 7) into each of two test tubes, labeled #3 and #4.
3. Add 5 drops of phenolphthalein to test tubes #1 and #3. Stir to mix thoroughly.
4. Add 0.1M sodium hydroxide, NaOH, solution dropwise to test tubes #1 and #3. Count the number of drops needed to change permanently the color of the liquid in each test tube. Record results.
5. Add 3 drops of methyl orange to test tubes #2 and #4. Mix thoroughly.
6. Add 0.1M hydrochloric acid, HCl, dropwise to test tubes #2 and #4. Count the number of drops of HCl needed to change permanently the color of the liquid in each test tube. Record results.

QUESTIONS:

1. Using information from Experiment 4-1 and the observations made in this investigation, explain why some of the salts in Part I produce a neutral solution when dissolved in water; why some produce an acidic solution; and why some produce a basic solution.
2. Using examples from your textbook, write hydrolysis equations to illustrate the action of the following salts with water: Na_2CO_3 and NH_4Cl .
3. Using Le Chatelier's Principle, explain:
 - a. the effect of adding a common ion to an equilibrium system composed largely of molecules.
 - b. the effect of adding a common ion to an equilibrium system composed largely of ions.
4. Assuming that 20 drops = 1 cc, then 5 cc of buffer contains 100 drops. Calculate what percent of the total volume of the buffer solution was required for the base in Part III to cause a permanent color change; for the acid to cause a permanent color change. How do these percentages compare with those of the distilled water?

Example: If it took 12 drops of base to cause the buffer to change color, then $12/100 = 0.12$. To change this to percent, we multiply 0.12 by 100% and get 12%. A volume of base equal to 12 percent of the total volume of the buffer solution was needed to cause a permanent change in the pH.

SUGGESTIONS FOR AN ADDITIONAL ACTIVITY

Devise an experiment to see whether various antacids and pain relievers contain buffering agents. Points to consider: why are we told to take these tablets with "a glass of water"? What do the advertisements tell us about these products? What contents are listed on the labels? Do they affect the normal pH of the stomach?

5-1 • BUILDING A MODEL FROM INDIRECT EVIDENCE

Because an atom cannot be observed directly, chemists must rely upon indirect evidence to describe its structure. The present concept of atomic structure has grown slowly. It is the result of evidence obtained from thousands of experiments. The diagrams we use to portray the atom are crude representations at best and must be considered as such.

Mental images and comparisons with familiar objects are ways used to develop models for science. The two tests of a model are (1) whether it explains what has been observed by experimentation and (2) whether it predicts what has not yet been observed.

To help you understand how a scientific model develops, you will be given a sealed box containing a number of objects. Your task is to determine by indirect methods what the box contains. You need also to arrive at a satisfactory model for the internal structure of the box.

PURPOSE:

To develop a reasonable model for the contents of a sealed container on the basis of indirect observations.

PROCEDURE:

1. Obtain a sealed container from your instructor.
2. By indirect methods, obtain evidence about the following:
 - a. The internal pattern of the box: Are there partitions? If so, how are they arranged?
 - b. The number of objects in the box.
 - c. The size, shape, mass, color, magnetic properties, material of each object.
3. You may conduct whatever tests you feel are necessary as long as you do not open the box.

CONCLUSIONS:

1. Sketch the internal pattern of the box based on the evidence you have collected.
2. Organize a table for all the observations you have made about the objects in the box. List the properties for each object, such as: size, mass, shape, composition, color, magnetic properties, etc.

QUESTIONS:

1. Why is an activity like this often called a "black box" experiment?
2. List some present-day "black boxes."
3. Are there properties that cannot be investigated indirectly?

5-2. IDENTIFICATION OF IONS BY COLOR

Certain ions can be identified by the color they give to a flame. These colors are related to the atomic structure of the elements. Identification of ions by flame tests is an important technique that you will use when you begin your study of qualitative chemistry—the “detective story” part of chemistry.

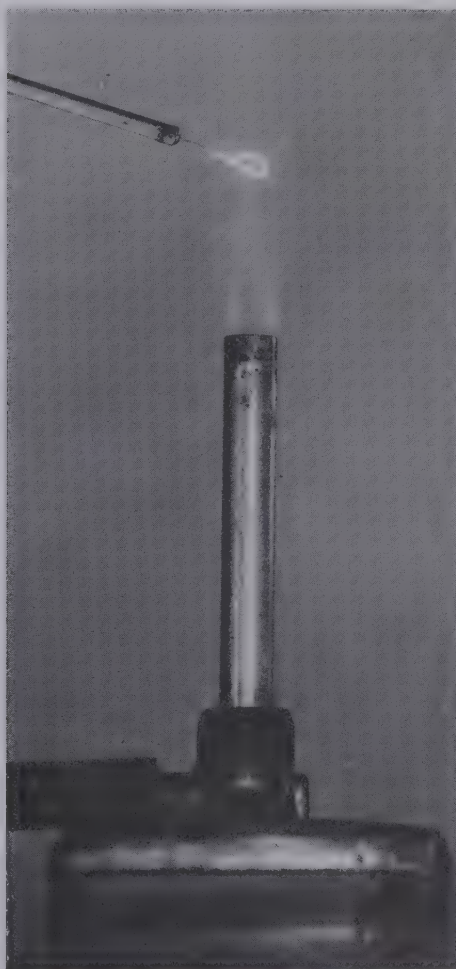
PURPOSE:

To observe the colors certain ions give to a flame and to develop skill in identifying “unknown” ions.

PROCEDURE:

Part I: Flame tests.

1. Obtain a platinum wire from the instructor. Pour a small amount of dilute HCl into a clean watch glass or test tube. Dip the platinum wire in the HCl and then heat it in the burner flame. This cleans the wire and insures that the observed flame color is produced by the new substance. As you dip the wire, be sure to keep the glass handle out of the liquid.





2. Pour 5 cc of each of the following solutions of ions into separate test tubes properly labeled:

sodium ions
potassium ions
calcium ions
barium ions
strontium ions
copper ions

Test each solution separately by dipping the tip of the wire into the solution and heating it in the flame. You may need to dip the wire into the solution several times to be sure of the flame color. (*Caution:* Clean the wire with HCl between solutions.) Record flame colors.

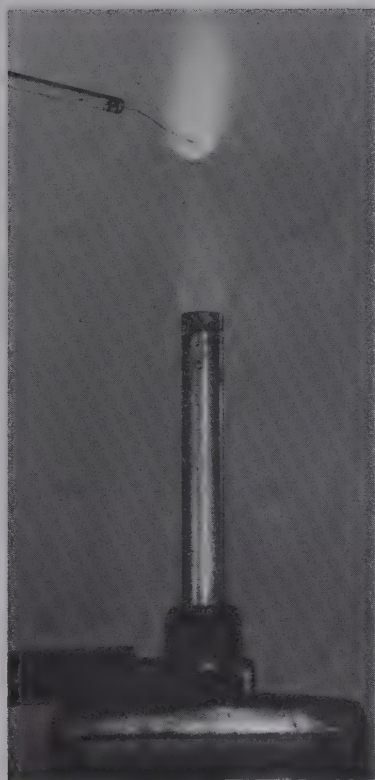
3. If a diffraction grating is available, view the flame colors through the grating. (Note: If possible, your instructor will set up a spectroscope so that you can view the hydrogen spectrum upon which so much of our evidence for atomic structure is based. If spectrum tubes of helium, argon, and mercury vapors can be obtained, you will find these interesting also.)

Part II: Borax Bead Tests

The borax bead test requires the use of the platinum wire and sodium tetraborate, commonly called borax.

1. Make a loop in the platinum wire by curling the tip around a pencil lead.
2. Clean the wire in HCl.
3. Pour about 2 cc of each of the following solutions of ions into separate test tubes:

cobalt ions
chromium ions
manganous ions
nickel ions

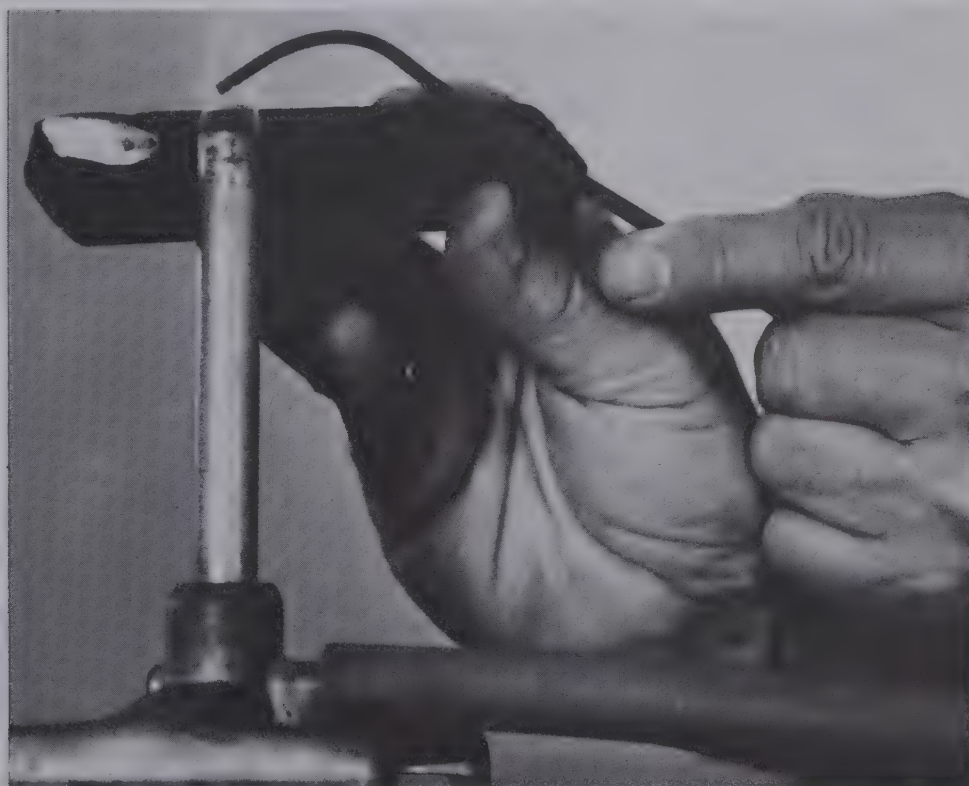


4. Place some powdered borax on a clean watch glass. Dip the loop of platinum wire into the borax and heat. You may need to dip the loop several times to form a clear, glass-like bead. (Note: If the bead is not perfectly clear, the wire was probably contaminated. Let the bead cool, crush it gently with a pencil, clean the wire, and begin again.)
5. When a clear bead has been formed, dip the hot bead into the first test solution and heat. Repeat this procedure until a definite color can be observed in the bead. Record the color in your data table.
6. Dispose of the bead and repeat the procedure with all solutions.

Part III: Cobalt Nitrate Tests

The cobalt nitrate test requires use of a charcoal block and cobalt nitrate solution.

1. Pour 5 cc of each of the following solutions of ions into separate test tubes:
 - aluminum ions
 - zinc ions
 - magnesium ions
2. Make a small depression in a charcoal block. Put several drops of one of the test solutions in the depression and heat strongly with a blowpipe in an oxidizing flame. (See illustration.) Remove the block from the flame. Add several drops of cobalt nitrate solution. Heat again with the blowpipe. Repeat these procedures until a definite color is seen on the charcoal block. Record observations.



3. Scrape the color off the charcoal block with a knife.
4. Repeat the above procedures with all solutions. Record observations.

QUESTIONS:

1. What evidence do you have that the color in each test was caused by the positive ion (metallic ion) and not by the negative (nonmetallic) ion?
2. Would you expect dry salts of these compounds to impart the same color to the flame, bead, or charcoal block? Why? Testing your answer would be a good additional lab activity.

SUGGESTIONS FOR ADDITIONAL ACTIVITIES:

Obtain samples of "unknown" substances and try to identify the positive ion present in each sample with the techniques of this experiment.

Investigate the composition of flares and fireworks to find the sources of their bright and varied colors. (*Caution:* This is to be a library activity only!)

5-3 • PAPER CHROMATOGRAPHY

Chromatography is an analysis technique for separating the parts of a mixture. It uses the process of selective adsorption. (Find out how the two terms *adsorb* and *absorb* differ.)

Solvents move through a porous material, like filter paper, by capillary action. The porous material is called the adsorbent. When a mixture is introduced into it, the adsorbent holds the different parts of the mixture with varying degrees of firmness. As the solvent moves through the adsorbent, it carries the most loosely held materials the greatest distance from the starting point. The parts of the mixture held more firmly by the adsorbent are not carried as rapidly nor as far. Thus the different parts of the mixture are separated.

A general relationship exists between the solvent and each part of the mixture. This relationship, called the rate of flow (R_f), can be expressed as:

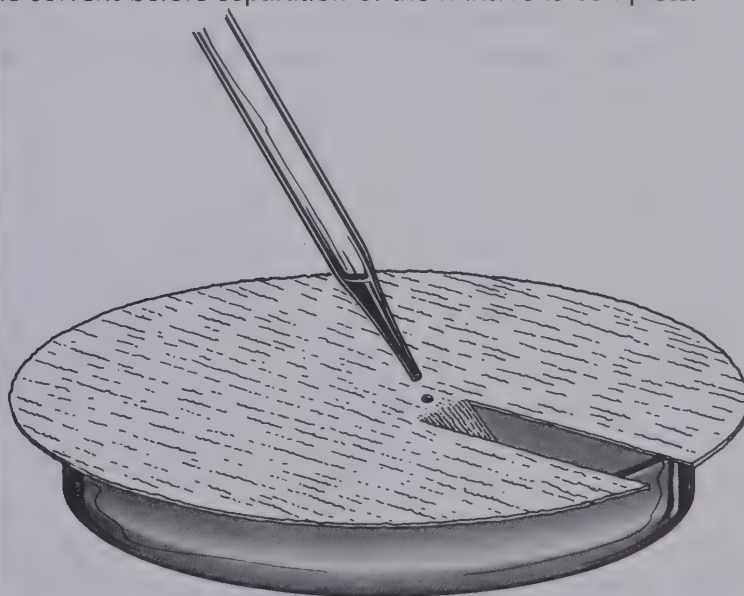
$$R_f = \frac{\text{distance traveled by the solute}}{\text{distance traveled by the solvent}}$$

PURPOSE:

To study the principles involved in chromatography by separating the components in ink.

PROCEDURE:

1. Use the separate halves of three Petri dishes as containers for the different solvents. Prepare pieces of filter paper as shown. Be sure that the tongue of the filter paper just barely dips into the solvent. The tongue should be just long enough to touch the bottom of the Petri dish. If the tongue is too long, or if too much solvent is placed in the dish, the filter paper is likely to collapse into the solvent before separation of the mixture is complete.



(Left) The tongue of the filter paper should barely touch the dish. (Right) Make the drop as small as you can.

2. Label filter papers along the outer edge *in pencil* with the solvent to be used. Use three solvents: water, ethyl alcohol, and ammonia water.
3. Barely cover the bottom of the Petri dish with the solvent.
4. Dip one end of tapered stirring rod into the ink and transfer a very tiny drop to the filter paper about 2 mm beyond where the tongue dips down.
5. Place the filter paper over the dish with the tongue barely touching the solvent. Observe results for about 15 minutes. Remove the filter paper from the dish before the solvent reaches the outer edge.
6. Using a pencil, carefully mark the outer edge of the solvent and the farthest movement of each pigment.
7. Calculate the R_f for each pigment. Make all measurements from the *top* of the original spot.
8. When the filter paper is dry, examine the chromatogram with a binocular microscope.

QUESTIONS:

1. In which solvent are the pigments most soluble? Least soluble?
2. Propose a hypothesis to account for any difference in the chromatograms produced by the different solvents.
3. Why should all measurements be made from the top of the ink spot?

SUGGESTIONS FOR ADDITIONAL ACTIVITIES:

Try different kinds of ink, even ball point.

Try separating the pigments in Universal Indicator, the juice of fresh spinach leaves, and other mixtures you can think of.

Read about column chromatography and design an experiment to separate a suitable mixture by this process.

Prepare a report on the use of gas chromatography in industrial processes.

6-1 • GROUPING ELEMENTS INTO FAMILIES ON THE BASIS OF SIMILAR ELECTRON PATTERNS

Now that you have studied the system by which we describe the electron pattern of an atom, you can use it to identify elements with similar structure. Elements can be grouped into chemical families according to similar chemical properties. The atoms of the elements in a family have the same number of electrons in the outermost energy level.

In this experiment you will explore the similarities in electron pattern of the first 54 elements by writing the electron configurations for them. Then by examining these configurations you will be able to identify those elements which belong together in a family. You will use the results of this experiment during the rest of your study of chemistry.

PURPOSE:

To group elements into chemical families on the basis of similar electron patterns.

PROCEDURE:

1. Figure out the electron configuration for elements 1 through 54, using the method described on p. 161 of the text. Be sure to rewrite each configuration so that all electrons in a given energy level are written together.

Example: $_{27}\text{Co}: 1s^2 \mid 2s^2 2p^6 \mid 3s^2 3p^6 3d^7 \mid 4s^2$

This is important because you will be interested in the *outermost* energy level as you group the elements.

2. Prepare eight vertical columns in your laboratory record book. Then look at the pattern of electrons in the outermost energy levels of the atoms of the elements for which you have written electron configurations. List in separate columns the symbols and atomic numbers of elements that have the following patterns of electrons in the outermost energy levels of their atoms.
 - a. A maximum number of electrons (Normally, this requires six electrons in *p* orbitals; what element is the exception to this rule?)
 - b. Only one electron in the *s* orbital
 - c. Two electrons only, both in the *s* orbital
 - d. Three electrons only
 - e. Four electrons only
 - f. Five electrons
 - g. Six electrons
 - h. Seven electrons
3. Prepare two horizontal rows in your record book long enough to contain the symbol and atomic number of ten elements. In the first row list all the elements which have no *4p* electrons but in which the next-to-the-outermost energy level contains electrons in *3d* orbitals. In the second row list

those which have no $5p$ electrons but in which the next-to-the-outermost energy level contains electrons in $4d$ orbitals. Cross out of your vertical columns any symbols and atomic numbers that you have entered in your horizontal rows.

QUESTIONS:

1. How many elements remain listed in each vertical column?
2. Look at the atomic numbers of the elements in each vertical column. What is the significance of the pattern? (Note: Take the numerical difference between the atomic numbers of consecutive elements in the column.)
3. How many elements are in each horizontal row? What is the significance of this number?
4. Write the electron configurations for the following elements: $_{55}\text{Cs}$, $_{57}\text{La}$, $_{82}\text{Pb}$, and $_{86}\text{Rn}$. Predict the vertical and/or horizontal rows where these elements would fit into the charts you have already made. Give a reason for your predictions.

6-2. THE RELATIONSHIP BETWEEN ATOMIC NUMBER AND THE RADIUS OF ATOMS

In building a model of the atom it is reasonable to expect that as the number of particles in the atoms increases, so should the size of the atoms. However, we must remember that all electrons are located around the nucleus and that electrons repel each other. What do experimental results tell us about the actual size of atoms?

Many methods have been used to compute and measure the size of atoms. Results vary with the method used. The atomic radius can be determined by spectroscopic methods, by x-ray diffraction, by measurements of the properties of gases, and in other ways. These techniques are too complicated to discuss here, but we can use the results gained from such experiments.

In this experiment you will graph the radii of atoms as a function of their atomic numbers. From this graph, you can make generalizations about the relationship of atomic number and atomic size.

PURPOSE:

To study the relationship between atomic number and atomic size to see whether there is a definite pattern.

Table of Atomic Radii in Angstrom Units*

1 H 0.30	15 P 1.10	29 Cu 1.28	42 Mo 1.36
2 He 0.93	16 S 1.04	30 Zn 1.33	43 Tc —
3 Li 1.52	17 Cl 0.99	31 Ga 1.22	44 Ru 1.33
4 Be 1.11	18 Ar 1.54	32 Ge 1.22	45 Rh 1.34
5 B 0.88	19 K 2.31	33 As 1.21	46 Pd 1.38
6 C 0.77	20 Ca 1.97	34 Se 1.17	47 Ag 1.44
7 N 0.70	21 Sc 1.60	35 Br 1.14	48 Cd 1.49
8 O 0.66	22 Ti 1.46	36 Kr 1.69	49 In 1.62
9 F 0.64	23 V 1.31	37 Rb 2.44	50 Sn 1.40
10 Ne 1.12	24 Cr 1.25	38 Sr 2.15	51 Sb 1.41
11 Na 1.86	25 Mn 1.29	39 Y 1.80	52 Te 1.37
12 Mg 1.60	26 Fe 1.26	40 Zr 1.57	53 I 1.33
13 Al 1.43	27 Co 1.25	41 Nb 1.43	54 Xe 1.90
14 Si 1.17	28 Ni 1.24		

*(Handbook of Chemistry and Physics, The Chemical Rubber Company, Cleveland, Ohio, 1966)

PROCEDURE:

1. Construct a graph with atomic numbers on the horizontal (x) axis and atomic radii on the vertical (y) axis. If you are not sure how to choose an appropriate scale, consult your instructor.
2. Join the points with a black line.

3. From your tabulated results from Experiment 6-1, determine which elements have the maximum number of electrons in the outermost energy level of their atoms. Divide the graph into sections by drawing vertical red lines through the atomic numbers of these elements.

TABULATION OF RESULTS:

1. List in separate vertical columns
 - a. the symbols and atomic numbers of the elements through which the red lines pass.
 - b. the symbols and atomic numbers of the elements with the smallest atomic radius in each section of your graph. The sections are marked by the red lines.
 - c. the symbols and atomic numbers of the elements with the largest atomic radius in each section of your graph.
2. List in a horizontal row
 - a. the symbol and atomic number of the elements in which 3*d* orbitals are filling.
 - b. the symbol and atomic number of the elements in which 4*d* orbitals are filling.

QUESTIONS:

1. Note the regular, repeating pattern in each section of the graph. How can you explain this pattern? Call upon your knowledge of electron configurations.
2. What can you say about the trend in atomic radius as the atomic number increases within a vertical column?
3. What can you say about the trend in atomic radius as the atomic number increases within each section of the graph? Give a reason for this trend.
4. What can you say about the trend in atomic radius as the atomic number increases while *d* orbitals are filling? Give a reason for this trend.

6-3 • THE RELATIONSHIP BETWEEN ATOMIC NUMBER AND IONIZATION ENERGY

In Experiments 6-1 and 6-2 we considered neutral atoms in their lowest energy state. Chemists have found that, by supplying sufficient energy, they can remove the most loosely held electron in an atom. The atom is thereby made a charged particle. Such a charged particle is called an ion. The energy required to remove the most loosely held electron is known as ionization energy. This energy is usually measured in electron-volts. We need not be concerned at this time with the definition of an electron-volt, other than to remember it is a definite quantity of energy.

You have discovered that the atomic number of an element is an important factor in relation to electron configuration and atomic radius. Are atomic number and ionization energy also related? In this experiment you will plot ionization energy versus atomic number in an attempt to answer this question.

PURPOSE:

To study the relationship of atomic number and ionization energy and to determine if a definite trend is present.

Table of Ionization Energies in Electron-Volts*

1 H 13.5	15 P 11.0	29 Cu 7.7	42 Mo 7.2
2 He 24.5	16 S 10.4	30 Zn 9.4	43 Tc —
3 Li 5.4	17 Cl 13.0	31 Ga 6.0	44 Ru 7.5
4 Be 9.3	18 Ar 15.8	32 Ge 8.1	45 Rh 7.7
5 B 8.3	19 K 4.4	33 As 10.5	46 Pd 8.3
6 C 11.2	20 Ca 6.1	34 Se 9.7	47 Ag 7.6
7 N 14.5	21 Sc 6.6	35 Br 11.8	48 Cd 9.0
8 O 13.6	22 Ti 6.8	36 Kr 14.0	49 In 5.8
9 F 17.3	23 V 6.7	37 Rb 4.2	50 Sn 7.3
10 Ne 21.5	24 Cr 6.8	38 Sr 5.7	51 Sb 8.6
11 Na 5.1	25 Mn 7.4	39 Y 6.6	52 Te 9.0
12 Mg 7.6	26 Fe 7.9	40 Zr 7.0	53 I 10.4
13 Al 6.0	27 Co 7.9	41 Nb 6.8	54 Xe 12.1
14 Si 8.1	28 Ni 7.6		

*(Handbook of Chemistry and Physics, The Chemical Rubber Company, Cleveland, Ohio, 1966)

PROCEDURE:

1. Plot a graph with atomic number on the horizontal (x) axis and ionization energy on the vertical (y) axis. If you are not sure of an appropriate scale, consult your instructor.
2. Join the points with a black line.
3. Divide the graph into sections by drawing vertical red lines through the atomic numbers of elements having the highest ionization energies. These are represented by the peaks on your graph.

TABULATION OF RESULTS:

1. List in separate vertical columns the elements having
 - a. the highest ionization energy in each section of the graph
 - b. the lowest ionization energy in each section of the graph
 - c. the next-to-highest ionization energy in each section of the graph (ignore the elements on the red lines)

CONCLUSIONS:

1. Describe the pattern formed by your graph. How can you explain the pattern? (NOTE: Relate the pattern to the electron configuration of the elements. Consider the filling of *s*, *p*, and *d* orbitals in your answer.)
2. What can you say about the effect of atomic number on ionization energy within a vertical column? Within a section of the graph (energy level)? Give a reason for your answer.
3. Place the graph of ionization energies on top of the graph of atomic radii and view them together against a strong light. What can you conclude about the trend shown by these two properties as plotted against atomic number? How can you explain this trend? (Hint: Locate the position of element 3, Li, on each graph. Compare the locations of other elements.)
4. Why are the same elements grouped together again and again when properties are plotted against atomic number?
5. Using the information obtained in Experiments 6-1, 6-2, and 6-3, make a generalization about the relationship of atomic number to the properties of elements.

6-4 • A CONVENIENT METHOD FOR GROUPING ELEMENTS

In Experiments 6-1, 6-2, and 6-3 you saw that when elements are arranged in order of increasing atomic number, similar properties recur at more or less regular intervals. Every eighth or eighteenth element has similar properties. You have also noted that elements with partially filled *d* orbitals have similar properties, different from the properties of elements in which *d* orbitals are not filling.

In this experiment you will make a chart in which the elements with similar properties are arranged in columns. This will be a convenient reference as you continue your study of chemistry.

PURPOSE:

To construct a chart in which elements with similar properties are grouped together.

PROCEDURE:

1. Mark a sheet of graph paper into squares about $\frac{1}{2}$ inch or 1.5 cm on a side. (The units will depend on the available graph paper.)
2. Using the information obtained in Experiments 6-1, 6-2, and 6-3, copy the vertical columns of elements onto the paper. In each square list the symbol and atomic number of one element.

H
1

3. When all elements previously listed in vertical columns have been copied, copy the elements listed in horizontal rows into two rows of squares. Again, use the symbol and atomic number.
4. Use a scissors to cut out the columns and rows so that you have whole strips for each column and each row.
5. Arrange the columns and rows in such a way that the atomic numbers read consecutively from left to right.
6. When you have completed the arrangement of the columns, paste the finished chart in your laboratory record book.
7. Label each column appropriately, such as, elements with 1 electron in outer energy level, elements with 2 electrons in outer energy level, etc.

QUESTIONS:

1. Knowing that elements may have electrons in *f* orbitals, and knowing that to date more than 100 elements have been found, consider these questions:
 - a. What information would you need to place element 110 on your chart?
 - b. Where would you place the elements whose *f* orbitals are filling with electrons?
2. List at least three advantages of having a chart like the one you have made.

6-5 • SEPARATION OF GROUP I CATIONS

When you have a mixture of several soluble compounds in aqueous solution, you can separate the ions in solution by a process called qualitative analysis. For the purpose of qualitative analysis the ions are grouped according to the differing solubilities of their compounds. Positive ions are called *cations*. Negative ions are called *anions*. The Group I cations include silver, Ag^+ , mercury (I), Hg_2^{2+} ; and lead, Pb^{2+} . This experiment will show you the way to separate and identify the cations in Group I. After you have done the experiment with a solution containing all three ions you will be given an "unknown" solution. Your problem then is to see which ions of the three are present.

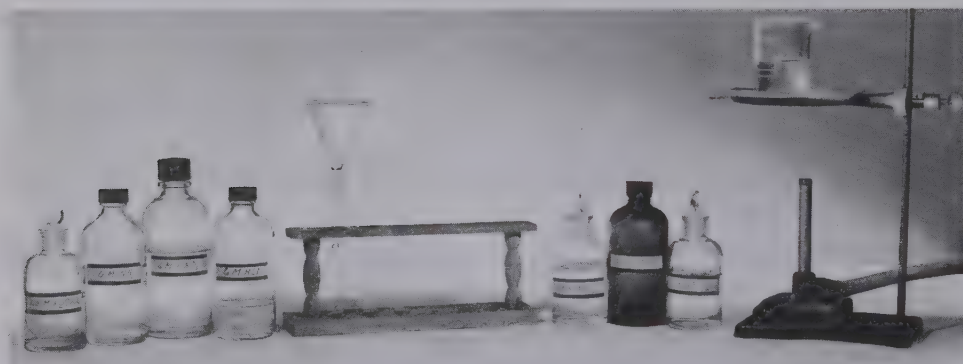
PURPOSE:

To become familiar with the ways to separate and identify cations in Group I.

PROCEDURE:

Part I: Separation and Identification of Ag^+ , Hg_2^{2+} and Pb^{2+} Ions

1. Pour 2 cc of each of the following solutions into the same test tube: silver nitrate, AgNO_3 ; lead nitrate, $\text{Pb}(\text{NO}_3)_2$; and mercury (I) nitrate, $\text{Hg}_2(\text{NO}_3)_2$.
2. Add 2 cc of HCl to the test tube. Mix thoroughly to permit complete precipitation.
3. Filter the contents of the test tube, catching the filtrate in a clean test tube. Rinse the filter paper and its contents with 5 cc cold distilled water. Discard filtrate.



The general array of solutions and equipment needed for Experiment 6-5.

4. Pour 5 cc of hot distilled water through the filter paper, catching the filtrate in a clean test tube. Set the funnel with the remaining precipitates over a clean test tube for further use. To the filtrate add 10 drops of potassium chromate, K_2CrO_4 , solution. A yellow precipitate identifies the Pb^{2+} ion.
5. Pour 2 cc of ammonia water through the filter paper. The black color on the filter paper identifies the Hg_2^{2+} ion.
6. To the filtrate from step 5 add nitric acid drop by drop until the solution is acidic to litmus paper. The appearance of a white precipitate identifies the Ag^+ ion.

Part II: Identification of the Cations in Group I in an "Unknown" Solution

1. Obtain a solution from the instructor which contains one or more of the Group I cations.
2. Determine which of the Group I cations are present in your solution by following the procedure learned in Part I. You should obtain evidence for the presence or absence of all three cations.
3. Report your findings to your instructor.

QUESTIONS:

1. The Group I cations are all precipitated by what anion?
2. Why is the group precipitate treated with hot water in Part I, step 4?
3. What causes the black color on the filter paper when the Hg_2^{2+} ion is treated with ammonia water?
4. Explain the function of nitric acid in the identification of the Ag^+ ion (Part I, step 6).

6-6 • IDENTIFICATION OF GROUP II CATIONS

Group II cations include antimony, Sb^{3+} ; bismuth, Bi^{2+} ; mercury (II), Hg^{2+} ; and copper (II), Cu^{2+} . These cations can be precipitated by hydrogen sulfide, H_2S , in a slightly acidic solution.

Hydrogen sulfide is a deadly poisonous gas and must be used with great care. When it is necessary to use this gas, prepare your solution to be tested before starting the H_2S generator. Then, as soon as possible after using the generator, flood it with water to prevent the escape of H_2S into the laboratory.

Hydrogen sulfide may be prepared by using a few small lumps of iron (II) sulfide, FeS , with HCl : $\text{FeS} + 2\text{HCl} \rightarrow \text{FeCl}_2 + \text{H}_2\text{S}$

An alternate method is by the use of thioacetamide. Your instructor will tell you which method you are to use in this experiment.

PURPOSE:

To become familiar with the technique for identifying cations in Group II.

PROCEDURE:

Part I: Identification of Sb^{3+} , Bi^{2+} , Hg^{2+} , and Cu^{2+} Cations

1. Place 2 cc of each of the following solutions in clean, labeled test tubes: copper (II) nitrate, $\text{Cu}(\text{NO}_3)_2$; antimony trichloride, SbCl_3 ; bismuth nitrate, $\text{Bi}(\text{NO}_3)_3$; and mercury (II) chloride, HgCl_2 . Add a few drops of 3M HCl to each solution.
2. Add H_2S or thioacetamide to each solution until a precipitate forms. Record the color of each precipitate in your data table.

Generating H_2S (Use this method only if thioacetamide is not available.)

- a. A 15 × 150 mm test tube fitted with a one-hole stopper and a delivery tube serves as a generator for H_2S . Always make sure you are ready to use the gas before starting the generator. H_2S is poisonous! If your laboratory has a fume hood, do this part of the experiment under the fume hood.



Bubbling H_2S from an H_2S generator (right) into a test solution. Remember to use the generator under a fume hood or at least in a very well ventilated location.

- b. When the test solutions have been prepared, put 3-4 small lumps of FeS in the generator and cover with HCl. Insert the stopper and bubble the gas into the first test solution.
- c. As soon as the test solution has time to react (usually 1-2 minutes), remove the delivery tube from the solution. Rinse the delivery tube and then allow H₂S to bubble into the second test solution. Repeat the procedure for the remaining test solutions. When all tubes have been treated with H₂S, quickly remove the stopper from the generator and flood the generator with water. Rinse the remaining FeS several times to remove acid, blot dry, and store for future use. Be sure to wash the delivery tube thoroughly to prevent contamination of the next test solution.

Part II: Identification of Group II Cations in an "Unknown" Solution

1. Obtain a solution from the instructor which contains one of the Group II cations.
2. Determine which Group II cations are present in your "unknown."
3. Report your findings to the instructor.

QUESTIONS:

1. What color is antimony sulfide? Copper (II) sulfide? Bismuth sulfide? Mercury (II) sulfide?
2. Assuming you have a solution containing both Group I and Group II cations, devise a way to separate and identify them.
3. What is the function of the hydrochloric acid in step 1?

7-1 • THE RELATIONSHIP BETWEEN BOND TYPES AND THE SOLUTION PROCESS

The process of solution is not completely understood at the present time. However, we do know that, for the solution to occur, the attractive forces between the solute particles must be overcome. If these forces are large, solution will not occur to any extent unless the attractive forces between the solute and solvent are also large. This means that the type of chemical bonds in both the solute and solvent are important in determining whether solution can occur. In this experiment, you will use polar and nonpolar solvents and polar, nonpolar, and ionic solutes. On the basis of your observations you will formulate principles about the effect of bond types on the solution process.

PURPOSE:

To determine the effect of bond types of both solute and solvent on the process of solution.

PROCEDURE:

Part I: Polar Solvent, Water (H_2O)

1. Place 5 cc of distilled water in each of six 13×100 mm test tubes. Label them A-1 through A-6.
2. Add 2 cc ethyl alcohol, $\text{C}_2\text{H}_5\text{OH}$, to test tube A-1.
Add 2 cc carbon tetrachloride, CCl_4 , to test tube A-2.
Add 2 cc oil to test tube A-3.
Add a few sugar crystals, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$, to test tube A-4.
Add a few crystals of sodium chloride, NaCl , to test tube A-5.
Add a small iodine crystal, I_2 , to test tube A-6.
(Caution: Use forceps.)



3. Place a small square of plastic over the top of each test tube and shake.
4. Observe the contents of each tube carefully to see whether the solute appears to have dissolved. Record results.

Part II: Nonpolar Solvent, Carbon Tetrachloride (CCl_4)

1. Place 5 cc of CCl_4 in each of six test tubes. Label them B-1 through B-6.
2. Add the same solutes as in Part I in the same order, except for B-2. Add 2 cc distilled water to B-2.
3. Place a square of plastic over each tube and shake.
4. Observe tube contents to determine whether solution has occurred. Record results.

Part III: Polar-Nonpolar Solvent, Ethyl Alcohol ($\text{C}_2\text{H}_5\text{OH}$)

1. Place 5 cc alcohol in each of six test tubes. Label them C-1 through C-6.
2. Add the same solutes as before in the same amounts, except for C-1. Add 2 cc distilled water to C-1.
3. Place a square of plastic over each tube and shake.
4. Observe contents of tube and determine whether solution has occurred. Record results.

CONCLUSIONS:

1. Use your knowledge of the periodic table, electronegativities, and bond types to classify each solute used in this experiment into these categories: ionic, polar covalent, and nonpolar covalent.
2. On the basis of this limited evidence, what can you say about the effect of bond types of solute and solvent upon solution?
3. How might these generalizations be of value in future experiments?

7-2. THE RELATIONSHIP BETWEEN BOND TYPES AND PHYSICAL PROPERTIES

The bonds in ionic solids are formed by a transfer of electrons between atoms. As a result, such solids consist of negative and positive ions arranged symmetrically. The bonds in molecular solids are formed by sharing electrons; ions do not exist in such solids. The forces between molecules in molecular solids are the relatively weak van der Waals forces.

In Experiment 7-1 you saw that bond types help to determine the solubility of a substance in a solvent. In this experiment you will investigate the relationship between bond types and the physical properties of solids.

PURPOSE:

To examine some of the physical properties of ionic and molecular solids and to relate these properties to the type of bond found in each solid.

PROCEDURE:

Part I: Hardness

1. Obtain samples of sodium chloride (an ionic solid) and paradichlorobenzene (a molecular solid).
2. Note the odor of each solid. An odor indicates that some of the solid's molecules can evaporate at room temperature.
3. Rub a sample of each solid between your fingers. Record the hardness of the solids by noting whether the "feel" is soft or granular.



Part II: Ease of Melting

1. Place a few crystals of sodium chloride in one tube and a similar amount of paradichlorobenzene in a second tube.
2. Heat each test tube separately. Heat cautiously at first and then more intensely if the solid does not melt readily. (Caution: If the solid does melt with gentle heating, do not heat it further. It may also vaporize, producing fumes. Be very careful not to inhale such fumes. They may be poisonous.)

Part III: Solubility

1. Place a few crystals of each solid in separate test tubes. Add water to each. Shake to mix. Compare the solubilities of the two solids in water.
2. Place a few crystals of each solid in test tubes containing benzene. Shake to mix. Compare the solubilities of the two solids in benzene.

SUMMARY OF OBSERVATIONS:

1. Prepare a chart in which you summarize the observations made on the ionic and molecular solids. This is part of your formal report.

QUESTIONS:

1. Using your knowledge of bond types and their relative strengths, explain your observations on the volatility of the two solids.
2. Name a volatile substance that does not have an odor.
3. How does bond strength affect the melting point of solids?
4. Refer to the results of Experiment 4-1 (Electrical Conductivity). Would you expect either or both liquid sodium chloride and liquid paradichlorobenzene to conduct an electric current? Give reasons for your answer.
5. Which melted solid would you expect to have the highest boiling temperature? Give reasons for your answer.
6. On the basis of evidence obtained in this experiment, make a general statement about the relative strength of the bonds in ionic solids as compared to those in molecular solids.
7. Using the periodic table and your electronegativity chart, predict whether each of the following substances is ionic or molecular:
 - a. sulfur
 - b. potassium iodide
 - c. calcium chloride
 - d. iodine
 - e. sulfur dioxide

7-3 • ENERGY IN CHEMICAL REACTIONS

All chemical reactions involve an exchange of energy. Bonds must be broken and new bonds formed. It takes energy to break bonds, and the formation of new bonds releases energy. The net effect of one bond breaking and a new one forming may be either an absorption or a release of energy. In the process of dissolving, which we often think of as being a physical change, energy is an important factor.

We classify chemical energy according to the kind of reaction with which it is associated—for example, heat of solution, heat of hydration, and so on. When equivalent amounts of an acid and a base react to produce a salt and water, the process is called neutralization. The energy associated with this process is the heat of neutralization.

In this experiment you will measure the amount of heat produced under three circumstances. The first is when solid sodium hydroxide, NaOH, dissolves in water. The energy released then is called the heat of solution. For the second you will measure the amount of energy released when equivalent solutions of NaOH and HCl are mixed. Third, you will measure the amount of heat released when solid NaOH neutralizes a solution of HCl. By comparing the amounts of heat from the three types of reactions we can learn a great deal about energy changes in different kinds of reactions.

The unit used to measure heat is known as the calorie. One *calorie* is defined as the amount of heat required to raise the temperature of 1 gram of water 1 Celsius degree. To simplify the calculations, we will ignore the heat absorbed by the plastic container, and assume that all heat is absorbed by the solution. The heat in calories can be obtained by multiplying the mass of the solution in grams by the change in temperature in Celsius degrees.

PURPOSE:

To determine the amount of heat evolved when solid NaOH dissolves in water (the heat of solution), when solutions of NaOH and HCl are neutralized (the heat of neutralization), and when solid NaOH reacts with HCl solution (the combined heats of solution and neutralization) and to examine the energy changes for a relationship among them.

PROCEDURE:

Caution: Be sure to wash and dry the cups between each part of the experiment. NaOH is corrosive. Do *not* handle it with bare fingers. Use forceps if needed. NaOH is also likely to become wet by absorbing moisture from the air, so work quickly and carefully. If a pellet falls on the table or floor while weighing, pick it up with forceps, place it in the sink, and run water on it until it dissolves.

Part I: Heat of Solution of NaOH

1. Pour 100 g of water into the plastic container.
2. Weigh out 1.0 gram of solid NaOH as quickly as you can.



When used as in this experiment, a plastic cup can be given the fancy name calorimeter.

3. Read and record the temperature of the water. Then add the pellets of NaOH to the water. Stir carefully with the thermometer and record the maximum temperature.
4. Calculate the number of calories of heat involved.

Part II: Heat of Neutralization of NaOH and HCl

1. Measure out 50 cc of 0.5M HCl into a plastic cup.
2. Measure out 50 cc of 0.5M NaOH into a second plastic cup.
3. Record the temperatures of both solutions, being careful *not* to transfer liquid from one cup to the other. Rinse and dry the thermometer between measurements. Both solutions should be at the same temperature and not above room temperature.
4. Pour the NaOH solution quickly into the HCl solution. Stir gently with the thermometer. Record the maximum temperature.
5. Calculate the number of calories of heat involved.

Part III: Combined Heat of Solution of NaOH and Heat of Neutralization of NaOH and HCl

1. Pour 100 cc of 0.25M HCl into a plastic cup.
2. Again weigh out 1.0 g of solid NaOH as quickly as you can.
3. Record the temperature of the HCl. Then add the NaOH pellets to the acid. Stir gently with the thermometer. Record the maximum temperature.
4. Calculate the number of calories of heat involved.

QUESTIONS:

1. If you add the calories of heat from reactions I and II, how does the total compare with the amount of heat evolved in reaction III? How can you account for this?
2. You used 1 gram of NaOH. How many moles is this?
3. Calculate the number of calories evolved per mole of NaOH in each reaction.
4. What is the source of heat in Part I? In Part II? In Part III?

8-1 • PREDICTING REACTION RATES

This experiment concerns the effect changes in the concentration of a base have on the time needed for the base to react completely with a metal.

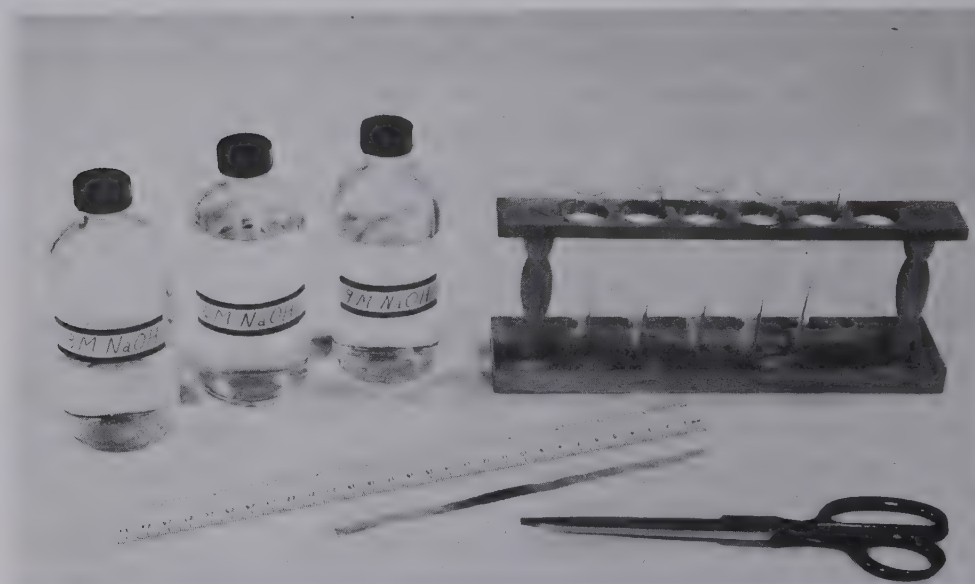
PURPOSE:

To observe the reaction between aluminum and sodium hydroxide at a given concentration and then to predict the time that will be required for other concentrations of the base to react completely with the same amount of aluminum.

PROCEDURE:

Part I: Reaction of Aluminum with 6M NaOH

1. Cut four 2 cm \times 0.5 cm strips of aluminum foil.



2. Pour 5 cc of 6M NaOH into a small test tube.
3. Decide how to get quantitative data for the reaction time.
4. Drop one strip of Al foil into the test tube. Record observations and the time required for the foil to be used up.

Part II: The Effect of Varying the Concentration of NaOH

1. Predict the reaction rates of Al with 3M, 6M, and 9M NaOH solutions. Record your predictions in your laboratory record book.
2. Test each concentration with one strip of Al foil as in Part I. Record your observations and the times required for the foil to be used up.

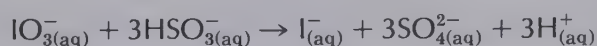
ANALYSIS OF DATA:

1. Graph the concentration of NaOH versus reaction times.
2. State a generalization which relates the two variables.
3. Propose hypotheses to explain the observed relationship between concentration and reaction rate.

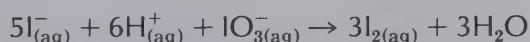
8-2. THE EFFECT OF CONCENTRATION AND TEMPERATURE ON REACTION RATE

The concentration of reactants and the temperature of a system are important in determining the rate of chemical reactions. We will investigate the effect of these factors in an interesting "clock" reaction that includes a noticeable rate-determining step. Two solutions are mixed and the time required to produce a color change indicates the rate of the reaction.

Solution A consists of potassium iodate, KIO_3 , in water solution and is the source of the iodate ion, IO_3^- . Solution B consists of sodium metabisulfite, $\text{Na}_2\text{S}_2\text{O}_5$, sulfuric acid, H_2SO_4 , and soluble starch. The sodium metabisulfite hydrolyzes to form sodium hydrogen sulfite, NaHSO_3 , in water solution. The initial step in the reaction is the slow step, or the rate-determining step:



When all the HSO_3^- ions have been consumed, the I^- ions react with the remaining IO_3^- ions to produce free iodine, I_2 :



In an equation the ions and molecules that are in water solution have the subscript (aq), from the Latin word *aqua* (water). The molecular iodine reacts with the starch to give a blue color.

PURPOSE:

To study the effect of concentration and temperature on the rate of a chemical reaction.



PROCEDURE:

Part I: The Effect of Concentration

Note: Your instructor may assign different dilutions to different pairs of students. By sharing the data, each pair can obtain a complete set of data without doing all the dilutions.

1. Label two small beakers A and B.
2. Pour 10 cc of Solution A into the first beaker and 10 cc of Solution B into the second beaker. Use a clean, dry graduated cylinder for each operation.
3. One student should time the reaction to the nearest second while a second student mixes the two solutions. Keep in mind that reaction starts when contact between the solutions occurs. So, start timing at the moment the solutions are mixed. When the timer is ready, pour Solution A into Solution B and quickly pour the mixture from one beaker to the other three or four times to insure complete mixing. The instant the solution changes color, record the time.
4. Prepare dilutions of Solution A as follows:

Solution A	Distilled Water
9 cc	1 cc
8 cc	2 cc
7 cc	3 cc
6 cc	4 cc
5 cc	5 cc
4 cc	6 cc
3 cc	7 cc
2 cc	8 cc
1 cc	9 cc

5. Perform as many reactions with dilutions of Solution A as assigned. Always mix the diluted Solution A with 10 cc of Solution B.

Part II: The Effect of Temperature

Note: Your instructor may assign different temperatures to various students.

For this portion of the experiment you will use 18 × 150 mm test tubes. A 250 cc beaker of water will serve as a water bath to heat the test tubes of solutions or to hold ice for cooling the solutions below room temperature.

1. Pour 10 cc of Solution A into one test tube and 10 cc of Solution B into the second test tube. Put both test tubes in a 250 cc beaker about $\frac{2}{3}$ full of water at the temperature you are assigned. Leave the solutions in the water (or ice) until they have come to the desired temperature.
2. When the solutions are at the desired temperature, pour them back and forth from one tube to the other three times until they are thoroughly mixed. One student times the reaction as in Part I. The instant a color change occurs, record the time for the reaction.
3. Repeat steps 1 and 2 for other assigned temperatures. Temperatures below 5°C and above 40°C give misleading results and should not be used. Above 40°C the HSO_3^- ions tend to decompose. Below 5°C the reaction is extremely slow.

QUESTIONS:

1. Why is it important to keep the total volume of Solution A at 10 cc during dilution?
2. Calculate the number of moles of KIO_3 in each cc of Solution A. The concentration of KIO_3 in the solution is 0.02M.
3. Find the number of moles of KIO_3 in each of the diluted solutions.
4. Plot a graph of class results on Part I. Plot time on the vertical axis and concentration of IO_3^- in moles/cc on the horizontal axis.
5. Plot a graph of class results on Part II. Plot time on the vertical axis and temperature on the horizontal axis.
6. On the basis of class results, what generalizations can you make about the effect of concentration on reaction time? How is reaction time related to reaction rate?
7. On the basis of class results, what can you say about the effect of temperature on reaction rate?

8-3 • TYPES OF CHEMICAL REACTIONS

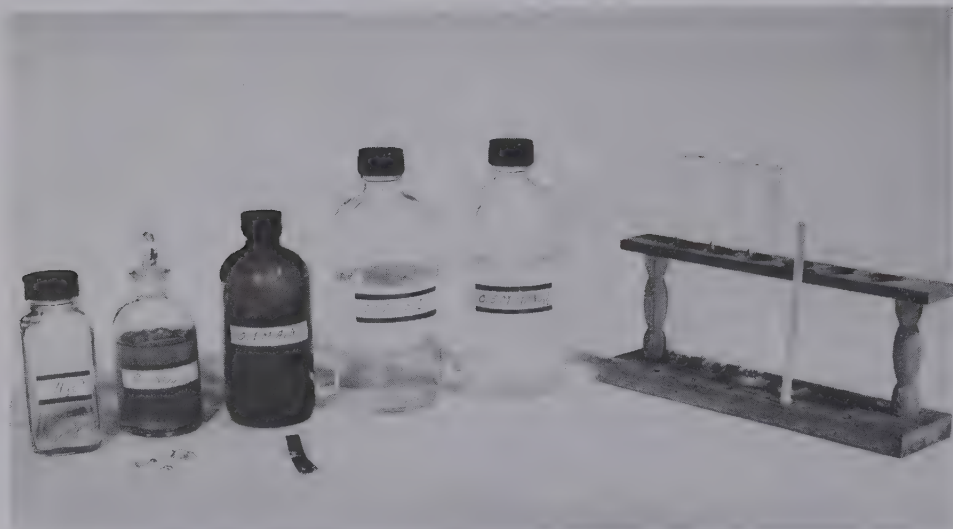
The most effective way to represent a chemical reaction is to write an equation. Now that you have studied electron configurations and the periodic table, it is time to develop skill in writing formulas and chemical equations.

In this experiment you will first observe examples of four major types of chemical reactions. You will then represent these reactions by written equations. It is suggested that you refer to all previous experiments in which chemical changes have occurred and write equations for those reactions.

The optional part of this experiment allows you to make quantitative calculations based upon chemical equations.

PURPOSE:

To observe chemical reactions and to practice writing chemical equations.



PROCEDURE:

Part I: Reactions Between Metals and Ions

1. Pour 4 cc of silver nitrate solution, AgNO_3 , into a test tube. Drop into the solution a 0.5 cm \times 3 cm strip of copper foil.
2. Pour 4 cc of copper sulfate solution, CuSO_4 , into a test tube. Add a small piece of mossy zinc.
3. Note whether changes occur immediately. Set these test tubes aside. Observe them at intervals as you proceed with Parts II, III, and IV.

Part II: Synthesis Reactions

1. Obtain a strip of magnesium ribbon about 4 cm in length. Note its physical properties.
2. Place an asbestos mat on the table and light your burner. Hold the magnesium ribbon by one end with the crucible tongs. Place the magnesium

ribbon in the flame until it ignites, then hold it over the asbestos mat until all reaction has stopped.

(Caution: Do not stare at the flame; it is a very intense light source.)

3. Note the physical properties of the product. Compare these with the properties of the magnesium ribbon.

Part III: Decomposition Reactions

1. Place a tiny pinch of mercuric oxide, HgO , in a small test tube. Note its physical properties.
2. Heat the mercuric oxide strongly for a few minutes. Test for oxygen gas by inserting a glowing (not burning) splint. When the reaction seems to be finished, remove the test tube from the flame. Note the properties of the residue. Consult your instructor about how to dispose of the residue.
3. Observe an electrolysis demonstration. Note the volumes of each of the gaseous products from time to time. Determine whether the voltage affects the rate of the reaction.

Part IV: A Reaction Involving Rearrangement of Ions

1. Pour 2 cc lead nitrate solution, $\text{Pb}(\text{NO}_3)_2$, into a test tube. Pour 2 cc of potassium iodide, KI , into a second test tube. Note the appearance of each solution.
2. Pour the lead nitrate solution into the potassium iodide solution. Record observations. (Note: Refer to Experiment 3-3, Part IV.)

EQUATIONS:

Complete the following word equations. Then write the correct formulas for each reactant and product. Finally, balance the equation using coefficients where needed.

1. copper + silver nitrate \rightarrow
2. zinc + copper sulfate \rightarrow
3. mercuric oxide \rightarrow
4. magnesium + oxygen \rightarrow
5. lead nitrate + potassium iodide \rightarrow
6. water \rightarrow
7. aluminum + hydrochloric acid \rightarrow
8. hydrogen + oxygen \rightarrow
9. aluminum chloride + ammonium hydroxide \rightarrow
10. iron + hydrochloric acid \rightarrow
11. iron sulfide + hydrochloric acid \rightarrow

QUESTIONS:

1. Classify each reaction above according to the four major types of chemical reactions. Give a reason for your classification.

2. List evidences of chemical action observed in Parts I, II, III, and IV of this experiment.
3. Refer to your notes on the following experiments and list evidences of chemical action observed during the reactions indicated:
 - a. Experiment 2-2: aluminum + hydrochloric acid \rightarrow
hydrogen + oxygen \rightarrow
aluminum chloride + ammonium hydroxide \rightarrow
 - Experiment 2-4: iron + hydrochloric acid \rightarrow
iron sulfide + hydrochloric acid \rightarrow

OPTIONAL:

1. In the reaction of copper with silver nitrate, assume that you wished to produce 10.8 grams of free silver. Show by use of the balanced equation how many grams of copper would be required. How many moles of silver is 10.8 grams? How many silver atoms does this represent?
2. If you burned 0.1 mole of magnesium ribbon, how many moles of magnesium oxide would be produced? How many grams would this be? How many molecules of magnesium oxide would be contained in this number of moles?

8-4. A QUANTITATIVE STUDY OF A CHEMICAL REACTION

In Experiments 3-3 and 8-3 you mixed potassium iodide with lead nitrate to produce lead iodide, a yellow precipitate. In this experiment you will do a quantitative study of the volumes of potassium iodide and lead nitrate necessary for a complete reaction. By a complete reaction, we mean one in which all of each reactant actually reacts.

Very careful measurements are important if your results are to be meaningful. Cleanliness is also essential. If either solution touches the other, some lead iodide forms.

PURPOSE:

To determine the volume ratio when equal strength solutions of potassium iodide and lead nitrate react completely and to explain your observations.

PROCEDURE:

Part I:

1. Put five 100 × 13 mm test tubes in a test tube support. Label them 1, 2, 3, 4, 5. Secure the labels as near the top of the tubes as possible. (Later in the experiment you will place the tubes in a beaker of water. If the labels get wet they may come off.)



2. Pour 4.0 cc of 0.5M potassium iodide solution into each test tube.
3. Using a clean graduated cylinder, add the following amounts of 0.5M lead nitrate solution to each test tube:

Tube	Potassium iodide	Lead nitrate
1	4 cc	0.5 cc
2	4 cc	1.0 cc
3	4 cc	2.0 cc
4	4 cc	3.0 cc
5	4 cc	4.0 cc

4. Place a plastic square over each test tube and shake to mix the solutions thoroughly.
5. Place the five test tubes in a 250 cc beaker. Carefully add water to the beaker until it comes just barely above the level of the liquid in tube 5. Be careful *not* to get any water into the tubes.
6. Heat the beaker until the water reaches the boiling point. Boil very gently for 2–3 minutes. Be careful that no water splatters into the tubes.
7. Remove the test tubes from the beaker. Stand them perfectly upright in a test tube support so that the precipitates will settle evenly. This is important in order to measure their heights accurately.
8. When the precipitates have settled, measure and record the height of each to the nearest millimeter.
9. Predict what adding *one* more drop of lead nitrate to each of the tubes will do. Will it change the height of the precipitate deposit in any of them? Record your predictions.
10. Add *one* drop of lead nitrate to each tube. Observe very closely to see if any more precipitate forms. Record your observations.

Part II:

1. Place five clean labelled test tubes in a test tube support.
2. Pour 2.0 cc of 0.5M lead nitrate into each test tube.
3. Using a *clean* graduated cylinder, add the following amounts of potassium iodide to the test tubes:

Tube	Lead nitrate	Potassium iodide
1	2 cc	0.5 cc
2	2 cc	1.0 cc
3	2 cc	2.0 cc
4	2 cc	3.0 cc
5	2 cc	4.0 cc

4. Place a plastic square over each tube and shake to mix the solutions thoroughly.
5. Heat the test tubes as you did in steps 5 and 6 of Part I.
6. Remove the tubes from the beaker. Set them aside so that the precipitates will settle evenly.
7. When the precipitates have settled, measure and record the height of each to the nearest millimeter.
8. Predict the effect of adding *one* more drop of potassium iodide to each tube. Will it change the height of the precipitate deposit in any of them? Record your predictions.
9. Add *one* more drop of potassium iodide to each tube. Observe very closely to see if any more precipitate forms. Record your observations.

ANALYSIS OF DATA:

1. Graph the height of the precipitates in Part I as a function of the volume of lead nitrate.

2. Graph the height of the precipitates in Part II as a function of the volume of potassium iodide.
3. What is the ratio of the volume of potassium iodide to the volume of lead nitrate needed for a complete reaction?
4. Write the word equation for the reaction between potassium iodide and lead nitrate.
5. Draw an electron dot diagram for each compound in your word equation. Determine whether the bond is ionic or covalent and show either transfer or sharing of electrons.
6. Propose a hypothesis to explain the ratio in your answer to question 3. Use your knowledge of bonding and your electron dot diagrams to help you.
7. In which test tube(s) in Part I were there apparently iodide ions that did not react? Give reasons for your answer.
8. In which test tube(s) in Part II were there apparently lead ions that did not react? Give reasons for your answer.
9. Apply Le Chatelier's Principle to explain your observations when one more drop of lead nitrate was added in Part I; when one more drop of potassium iodide was added in Part II.
10. Using all the information obtained in this experiment, how many moles of potassium iodide do you think are required to react completely with 1 mole of lead nitrate?
11. Explain your answer to question 10 by referring to the electron dot diagrams and by the evidence obtained in this experiment.
12. Write the balanced chemical equation for the reaction between potassium iodide and lead nitrate.

9-1 • DIFFUSION OF GASES

If the air is pumped from a container and a gas is then introduced, the container instantly fills with the gas. If gaseous molecules enter a container already occupied by other gases, the spreading is slower because of the frequent collisions. However, in time the gases become uniformly mixed because of their random motion. The process of spreading out spontaneously in all directions is called diffusion.

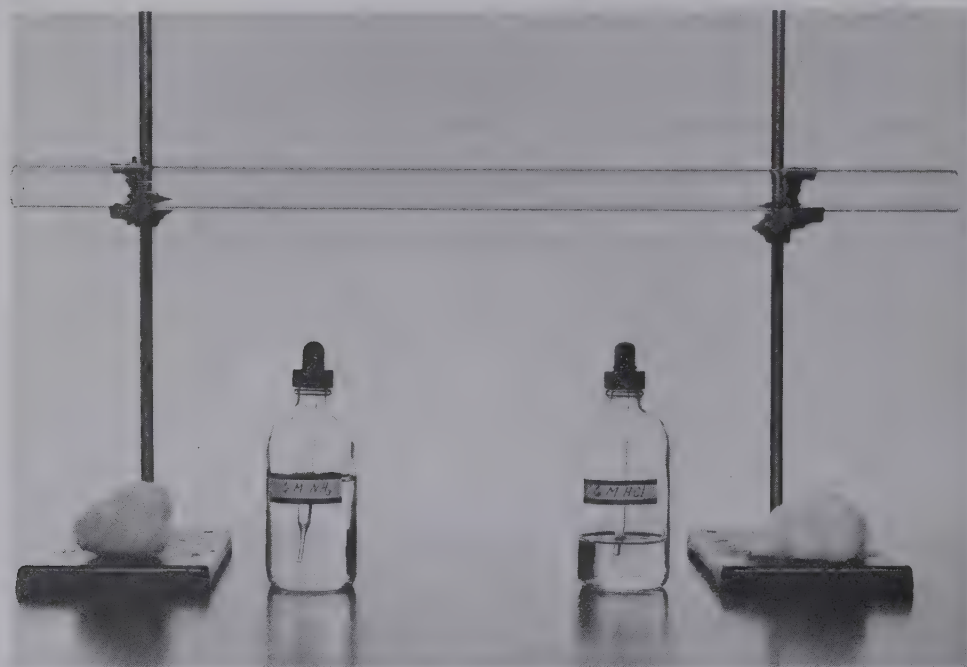
All gases at the same temperature have the same average kinetic energies. Do they have the same average speeds? Kinetic energy is dependent upon both the speed of the particle and its mass or weight. In this experiment you will introduce two gases of different molecular weights into a tube of air to study their rates of diffusion. You can assume they are at the same temperature. Why?

PURPOSE:

To study the diffusion rates of two gases.

PROCEDURE:

1. Set up apparatus as shown in the photo. Be sure the tube is level.



2. Secure two cotton plugs.
3. Using dropper pipets, place 6 drops of concentrated HCl on one plug and 6 drops of ammonia water on the other. Try to insert the plugs into opposite ends of the tube at the same instant. You may need a partner to help. (Note: Keep the bottles of reagents as far apart as possible.)
4. Observe the tube for 8–10 minutes. Record your observations. The appearance of a white ring marks the meeting place of the gases.
5. Measure and record the distance traveled by each gas, HCl and NH_3 .

6. If time permits, wash and dry the tube and repeat the experiment. If you cannot do this, perhaps you can obtain distance data from other members of the class. Frequently, a class average provides better data than one trial.
7. Clean and dry the tube and return it to its proper place.

QUESTIONS:

1. What is the composition of the white ring? What type of reaction is this?
2. How can you account for a series of white rings instead of a single white ring after the tube has stood for a long time?
3. Does the ring form progressively nearer the HCl end of the tube or the NH_3 end? How do you account for your observation?

SUGGESTIONS FOR ADDITIONAL ACTIVITIES

Vary the amounts of ammonia water and concentrated HCl used.

Vary the position of the glass tube. Make one end slightly higher than the other, or much higher, or make the tube vertical. Change the relative positions of the two liquids at each end.

9-2 • THE EFFECT OF PRESSURE ON THE VOLUME OF GASES

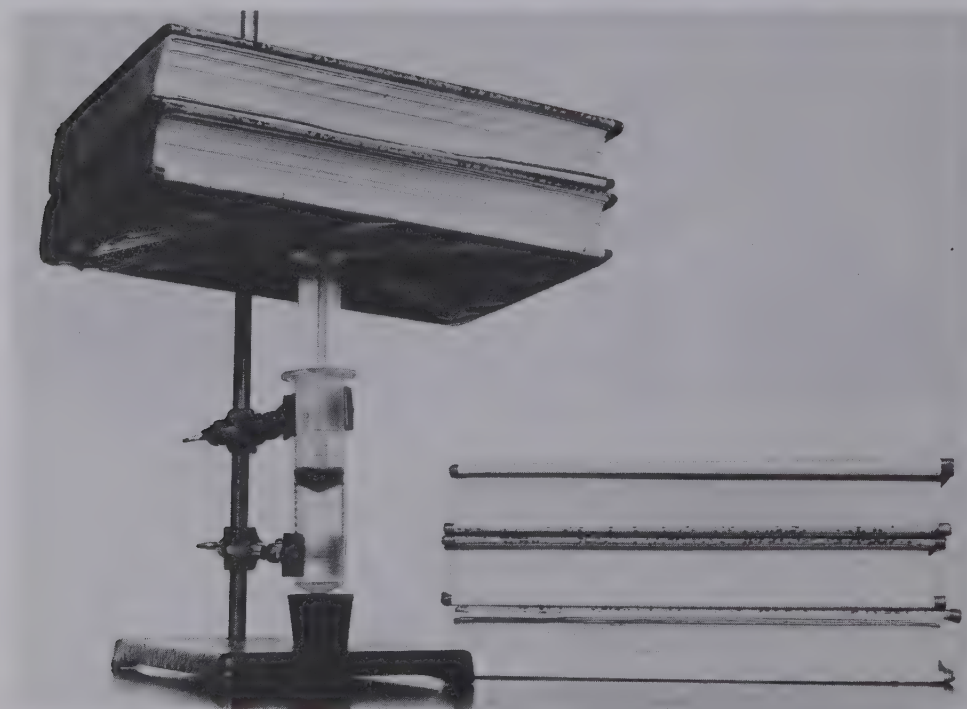
Automobile tires must hold air at a pressure well above that of the atmosphere if they are to support the car. When a tire is punctured, the air inside the tire escapes until the inside pressure equals that of the atmosphere. In this experiment you will study the effect of changing pressure on the volume of gases. The temperature of the gas will remain constant.

PURPOSE:

To study the effect of pressure on the volume of gases at constant temperature and to establish a relationship between pressure and volume.

PROCEDURE:

1. Obtain a small plastic syringe. Pull the plunger out to the maximum scale reading of the syringe. Plug the tip of the syringe and try to push the plunger in. Note the pressure as the air "pushes back."
2. Remove the plug and push the plunger in as far as possible. Place your finger over the tip as you pull the plunger out. Record observations.
3. Pull the plunger out to the maximum scale reading. Obtain a specially prepared rubber stopper from your instructor. Fit the tip of the syringe into this stopper, making sure it is in tight so the air cannot escape as the plunger is pushed in.
4. Mount the apparatus as shown in the illustration.
5. Obtain 5-6 copies of the same book or 5-6 other identical objects that can be used as illustrated. Place one book on top of the plunger. Read and record the volume of air trapped in the syringe to the nearest cc.



6. Continue to add books, one at a time. Read the volume of the gas each time a book is added.
7. Repeat steps 5 and 6 several times to check your readings.

ANALYSIS OF DATA:

1. Take an average of the gas volumes for *each* pressure. That is, average volume for a pressure of one book; average volume for a pressure of two books, etc.
2. Plot a graph of the average volume in cc on the vertical axis versus pressure in books on the horizontal axis. Join the points with a smooth curve.
3. For each point on your graph, multiply the average volume in cc by the pressure in books and compare the products.
4. On the basis of your observations, state a generalization about the effect of pressure on the volume of gases at constant temperature.
5. How does pressure affect the density of gases? Support your answer.
6. If you wished to double the volume of a gas at constant temperature, how would you change the pressure?

OPTIONAL QUESTIONS:

1. Without books on the plunger, what is the pressure on the sample of gas?
2. How many books are needed to reduce the volume by one half?
3. What is the value of atmospheric pressure in books?

9-3. THE EFFECT OF TEMPERATURE ON THE VOLUME OF A GAS

You may have observed that heating a balloon containing air causes the balloon to expand. This indicates that as the temperature of gases rise their volume must increase. In this experiment you will measure a gas volume at temperatures ranging from 0°C to near 100°C to determine whether there is a predictable relationship between temperature and volume.

PURPOSE:

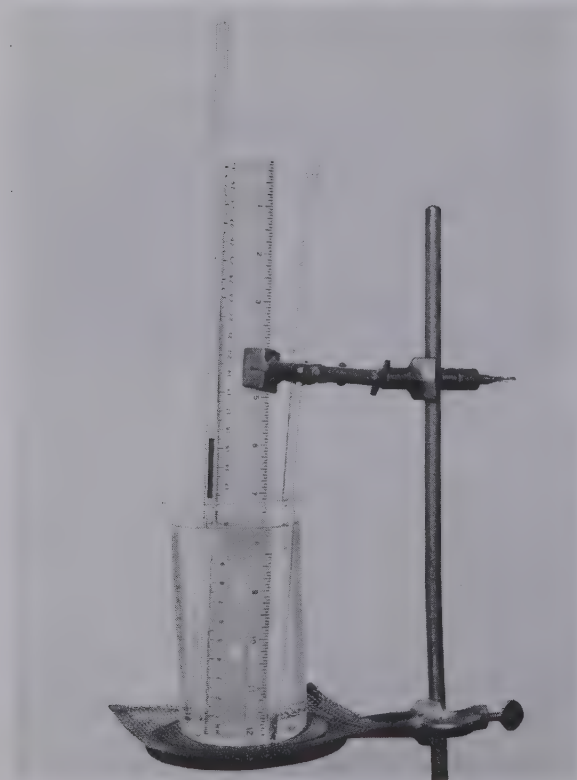
To study the effect of temperature on the volume of a gas and to try to establish a relationship between them.

PRECAUTIONS:

1. The apparatus you will use contains a drop of mercury. The tube is open at one end, so be careful not to spill the mercury. Since it fuses with other metals, take off all jewelry before doing the experiment.
2. Precision in reading the gas volumes is very important to good experimental results. Recheck each reading before going on to the next.
3. Precise thermometer readings are also important.

PROCEDURE:

1. Set up the apparatus as shown in the illustration, but leave the beaker empty. Be sure the scale is plainly visible at all points to insure precise readings throughout the experiment.



The gas sample in the tube should be surrounded by the water bath as fully as possible for it to be at the same temperature as the water.

2. Record the volume of trapped air to the nearest 0.1 cm. This is the volume at room temperature.
3. Fill the tall 400 cc beaker almost full of an ice-water mixture. The temperature should be at 0°C. The trapped air should be almost completely surrounded by the ice-water bath.
4. Wait 3 to 4 minutes until the air has had time to cool. Read and record the new volume.
5. Remove the beaker of ice-water. Replace it with a beaker of water some 15° or 20° warmer. Repeat the above procedure. Record both the water temperature and the air volume.
6. Continue the procedure until 5 or 6 readings have been taken to near 100°C.
7. If time permits, repeat the experiment to obtain another set of readings.

ANALYSIS OF DATA:

1. Graph the data. Can you determine the mathematical relationship between temperature and volume?
2. Extrapolate the line to temperature values below 0°C. At what Celsius temperature does the line cross the temperature axis? What is the significance of this temperature?

9-4 THE MOLAR VOLUME OF A GAS

The molar volume of a gas is the space occupied by 1 mole of the gas at standard temperature and pressure (S.T.P.). In this experiment you will produce hydrogen gas by reacting magnesium with hydrochloric acid. We know that in this reaction, one mole of magnesium produces one mole of hydrogen gas. By using a known weight of magnesium and by measuring the volume of hydrogen produced, you can find the volume one mole would occupy at S.T.P. The volume of hydrogen collected at room temperature and pressure needs to be corrected to S.T.P. Corrections need also be made for the water vapor which mixes with the hydrogen.

PURPOSE:

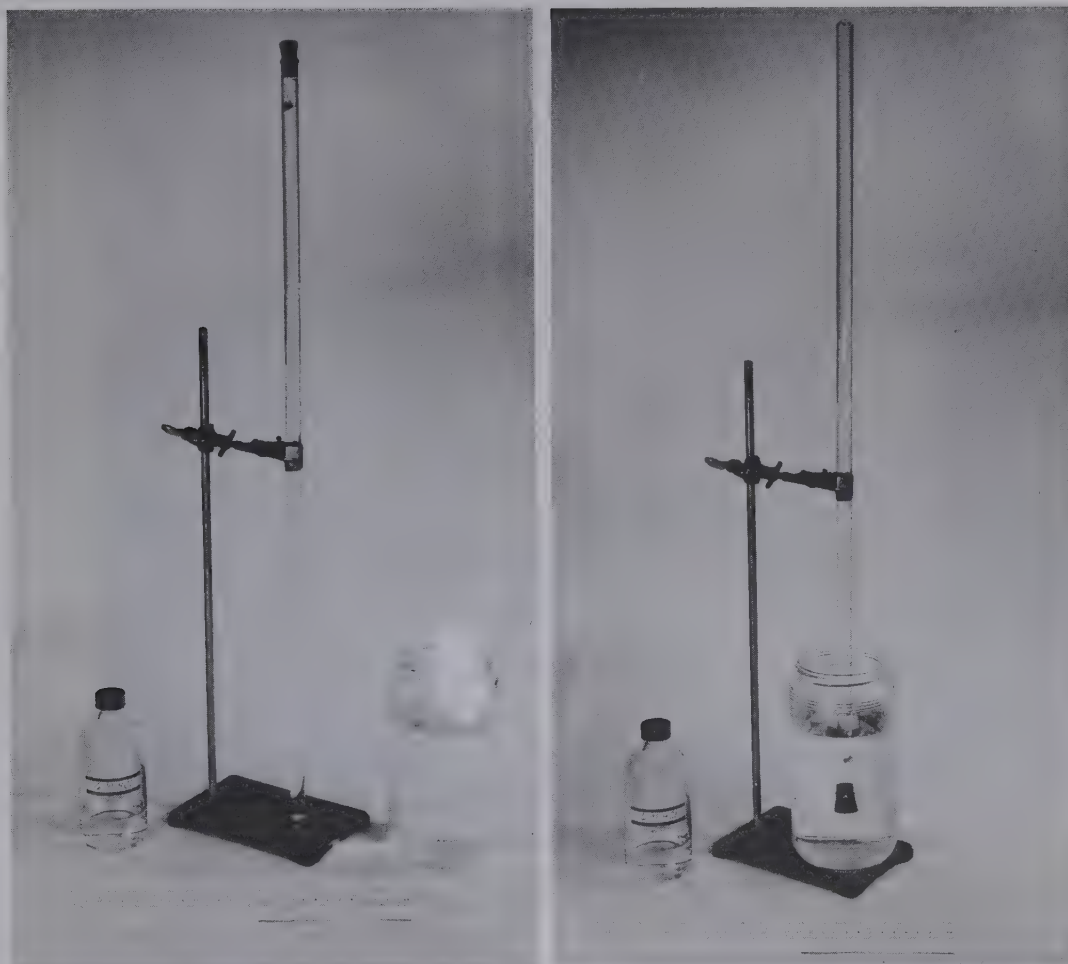
To find the volume occupied by 1 mole of hydrogen gas at S.T.P.

PRECAUTIONS:

1. Be careful not to spill the concentrated acid. Flush the area with water if spillage occurs.

PROCEDURE:

1. Obtain a pre-cut strip of magnesium ribbon from your instructor. Measure it very carefully to the nearest 0.05 cm and record the length. Your instructor will give you the weight of 1 meter of magnesium ribbon. From this you can find the weight of your strip.
2. Fold the magnesium ribbon loosely and enclose it in a cage of fine copper wire. Leave about 5 cm of the wire as a tail.
3. Set up a ringstand and attach a burette clamp as shown in the illustration.
4. Fill a large wide-mouth jar about $\frac{2}{3}$ full of water.
5. Pour 10 cc of 6M hydrochloric acid into the gas collecting tube. Carefully pour enough water down the inside of the tube to fill it to overflowing. Do not disturb the acid in the bottom of the tube any more than necessary.
6. Tap the tube to dislodge any air bubbles that may be clinging to the sides.
7. Insert the coil of magnesium into the tube to a depth of about 5 cm. Let the tail of copper wire project out the end of the tube. It will be held in place by the one-hole stopper.
8. Cover the hole in the stopper with your finger and invert the tube into the jar of water. Be careful not to let air into the tube.
9. As the acid mixes with the water the reaction will begin. When the reaction is complete, wait about 10 minutes until all hydrogen bubbles have left the water. Tap the tube gently with a pencil to help dislodge bubbles. It is very important to measure the total volume of hydrogen gas as precisely as possible.
10. Raise or lower the tube until the water level inside the tube coincides



The photo on the left shows the tube before it is inverted; the photo on the right shows it after it is inverted.

with the level of the water in the jar. Read and record the volume of hydrogen gas to the nearest 0.05 cc.

11. Measure and record the temperature of the water in the beaker.
12. Read and record the barometric pressure.

DATA AND CALCULATIONS:

1. Length of magnesium ribbon in meters (measured).
2. Weight of magnesium ribbon (calculated).
3. Number of moles of magnesium (calculated).
4. Barometric pressure (measured).
5. Partial pressure due to water vapor in tube (read from table).
6. Partial pressure of hydrogen in tube (calculated).
7. Volume of hydrogen corrected to S.T.P. (calculated).
8. Volume of hydrogen produced by 1 mole of magnesium (calculated).



The volume of the box, 22.4 liters, equals the volume of a mole of gas at 0°C and 760 torr. How much larger must the box be for it to hold a mole of gas at room temperature, at the same pressure?

9. Experimental error (calculated).
10. Percentage error (calculated).

QUESTIONS:

1. One mole of hydrogen gas weighs 2 grams. What is its density at S.T.P.?
2. Why does 1 mole of magnesium produce 1 mole of hydrogen gas from hydrochloric acid?

9-5 • TESTING PREDICTIONS IN THE LABORATORY

In Experiment 9-4 you used magnesium and hydrochloric acid to produce hydrogen gas. In this experiment you will make certain predictions based upon the observations and results from Experiment 9-4.

PURPOSE:

To predict the weight of magnesium or zinc that would be needed to react with HCl to produce a given volume of hydrogen. (You are free to say what volume you wish to produce.)

PROCEDURE:

1. Decide upon the volume of hydrogen you wish to produce. Write the balanced chemical equation for the reaction. Predict the number of grams of magnesium or zinc you will need to produce your volume of hydrogen.
2. Set up the appropriate apparatus. Obtain the weight of metal you have predicted.
3. Check the volume of hydrogen obtained against your predictions.
4. Repeat these steps, using different weights of zinc or magnesium until you have at least three sets of data for both metals.

ANALYSIS OF DATA:

1. Calculate the number of moles of hydrogen produced for each mole of zinc and magnesium used.
2. For each metal graph the number of moles of hydrogen produced versus the weight of the metal used.

QUESTIONS:

1. For each metal what relationship seems to exist between the weight of the metal and the number of moles of hydrogen produced upon reaction with HCl?
2. Propose a hypothesis to explain the relationship stated in question 1.

OPTIONAL:

1. Repeat steps 1-3 using aluminum metal. Compare the number of moles of hydrogen produced using Al with the number of moles produced with Zn and Mg. Relate these findings to the atomic weights of the three metals.
2. Propose a generalization which relates number of moles of hydrogen produced to the atomic weights of the metals used.

9-6 • DETERMINATION OF THE FORMULA FOR AN UNKNOWN COMPOUND

In Experiment 2-4, you calculated the percentage composition of iron sulfide which you had prepared. From this you found the simplest formula for this compound. You have also learned that 1 mole of any gas occupies 22.4 liters at S.T.P. In this experiment you will vaporize an unknown compound, find its molecular weight, and finally, determine the molecular formula for the substance. Your instructor will give you the percentage composition.

PURPOSE:

To vaporize an unknown compound and determine its molecular weight and its molecular formula.

PROCEDURE:

1. Obtain a clean, dry 150 ml Erlenmeyer flask.
2. Cut a piece of aluminum foil about 5 cm square. Weigh the dry flask and foil together to the nearest 0.01 g. Record the weight.
3. Place about 5 ml of the unknown liquid in the flask. Cover the flask with the aluminum foil. Crimp the foil tightly about the neck of the flask, rolling it up on the rim of the flask to make a tight seal.
4. Use the point of a pin to prick a very tiny hole in the center of the aluminum foil. The hole should be as small as possible.
5. Place a 1 liter beaker on an asbestos mat over the burner as shown in the photo.
6. Attach a burette clamp to a ringstand and support the flask by the neck. Do not disturb the foil.
7. Lower the flask into the beaker as shown. Attach the burette clamp firmly to the rod.



The flask should be as deep in the water as you can get it for its temperature to be the same as that of the water bath.

8. Carefully fill the beaker with water so that as much of the flask is immersed as possible.
9. Light the burner. Heat the water to boiling. The unknown liquid will probably be completely vaporized before the water comes to a boil. By rocking the flask ever so gently from time to time, you can see whether any liquid remains in it. Boil the water for a minute or so even though the liquid in the flask has already vaporized.
10. While the water is boiling, insert the thermometer into the beaker of boiling water about half way down the side of the flask. Do not let the thermometer bulb touch either the glass flask or the beaker as you read the boiling temperature of the water.
11. Turn off the burner. Carefully remove the clamp and flask. Use a hot pad so you do not burn your hands.
12. Let the flask cool by setting it on a piece of paper or cloth toweling on the lab bench. Remove the burette clamp. When cool, dry the flask thoroughly.
13. Without removing the foil, weigh the flask with its contents. Record the weight.
14. Use a graduated cylinder to measure the water needed to fill the flask level full. This represents the volume of the vapor of the unknown liquid when the water outside the flask was boiling.
15. Record the barometer reading.
16. Obtain the percentage composition of your unknown from your instructor.

DATA AND CALCULATIONS:

1. The barometric pressure in torr (measured).
2. Weight of flask, foil, and contents at end of experiment (measured).
3. Weight of empty flask and foil cover (measured).
4. Weight of unknown at end of experiment (calculated).
5. Volume of flask in cc (measured).
6. Volume of flask in liters (calculated).
7. Temperature of boiling water (measured).
8. Volume of vapor in liters at S.T.P. (calculated).
9. Molecular weight of unknown compound (calculated).
10. Percentage composition of unknown compound (obtained from instructor).
11. Simplest formula of unknown compound (calculated).
12. Molecular formula of unknown compound (calculated).

QUESTIONS:

1. Why was a tiny hole made in the aluminum foil cover to the flask?
2. What are the major sources of experimental error?

3. Why was the boiling temperature of the water used in correcting the vapor volume to S.T.P. instead of the room temperature?
4. Why did all the vapor of the unknown liquid not escape from the flask?
5. What was the pressure of the vapor after all of it that would escape had escaped?

RELATED QUESTIONS:

1. During the heating of the water, many bubbles collected on the sides of the beaker. Where did they come from? What was their composition?
2. As the water began to boil, why did most of the bubbles seem to rise from the same points at the bottom of the beaker?

10-1 • WATER IN CRYSTALS

Some crystalline solids contain a definite amount of water as a part of the crystal structure. This is known as the water of hydration. Upon exposure to air, hydrated crystals can lose their water of hydration if their vapor pressure is greater than that in the air around it. This process is called efflorescence.

Other solids may remove water vapor from the air to form a solution on their surfaces. This happens when the water vapor pressure of the air is greater than that of the solution. Under these conditions, the solid may absorb enough water to dissolve completely. This process is called deliquescence.

A few crystals hold minute traces of water in the crystal mechanically rather than in hydrated form. That is, the water is not an essential part of the crystal structure. When such solids are heated, the water inside vaporizes and fractures the crystal as it expands. This process is known as decrepitation.

PURPOSE:

To study a number of crystalline solids and to determine which ones contain water of hydration.

PROCEDURE:

Part I: Efflorescence and Deliquescence

1. Label four watch glasses and place about one teaspoonful of the following solids on separate ones: sodium carbonate, sodium hydroxide (*caution: use forceps, do not handle*), sodium chloride, calcium chloride.
2. Weigh each watch glass with its contents to the nearest 0.01 g. Record weights in data table.
3. Set the watch glasses aside until the following laboratory period. At that time, examine the contents of the watch glasses and record any changes. Again weigh each watch glass with its contents. Record weights in the data table. Classify each solid as to whether it is efflorescent, deliquescent, or neither.

Part II: Decrepitation

1. Place a few crystals of sodium chloride in a small test tube. Heat the test tube gently over a flame. Observe closely for evidence of decrepitation. When the test tube has cooled, examine the crystals with a magnifier.
2. Test potassium chloride and sodium chromate in the same manner.

Part III: Water of Hydration

1. Put a small crystal of copper (II) sulfate in a small test tube. Holding the test tube horizontally, heat the crystal gently with a burner flame. Observe any changes during heating.
2. When the test tube has cooled, pour the contents out onto a clean, dry, watch glass. Compare the form and color with an unheated crystal of copper (II) sulfate. Record observations.

3. Add several drops of water to the material that was heated. Note results. Check for production of heat.

QUESTIONS:

1. Give a practical reason for keeping the lids on reagent bottles of solids.
2. Name one possible commercial use of deliquescent solids.
3. What seems to cause the blue color of copper (II) sulfate?
4. How might atmospheric conditions affect the results in this experiment?
5. Consult a reference book to learn why sodium chloride collects moisture on a rainy day. You may be surprised.

OPTIONAL:

1. Devise a way to find the percentage of water in hydrated copper (II) sulfate.

10-2 • THE STRUCTURE OF CRYSTALS

Most solids exist in crystalline form. Their particles are arranged in a regular pattern throughout the solid. Such solids may be classified according to the kind of bonds that hold them together. On this basis there are four major types of crystals: (1) ionic, (2) molecular, (3) network, and (4) metallic. Sodium chloride is an example of an ionic crystal. Ice, sulfur, and iodine are molecular crystals. Diamond is a network solid, and silver and copper are metallic. In this experiment you will study examples of ionic, molecular, and metallic crystals.

PURPOSE:

To study the different types of crystal structure.

PROCEDURE:

Part I: Ionic Crystals

1. Label four glass slides, #1, #2, #3, and #4. Place one drop of the following saturated solutions on the given slides.

#1 - sodium chloride

#2 - magnesium sulfate

#3 - nickel sulfate

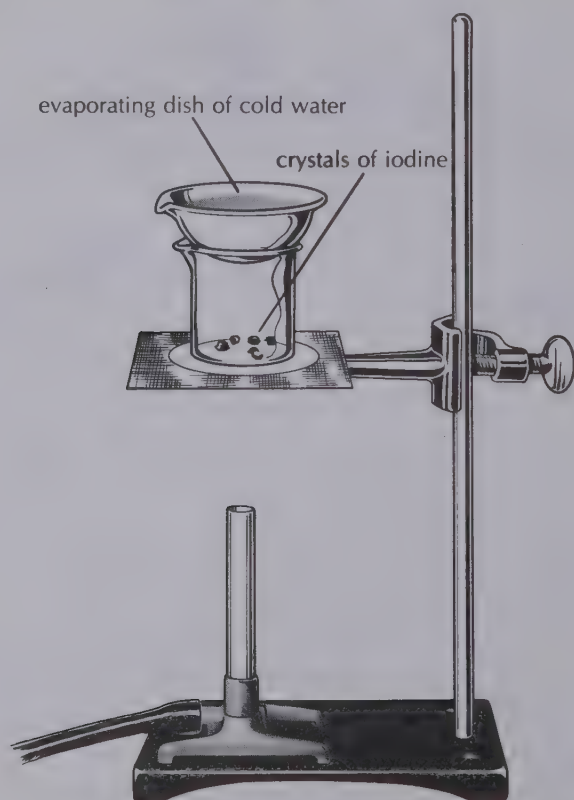
#4 - potassium nitrate



2. Evaporate the liquid from each slide by gently warming it over a flame. Be careful to hold the slide well above the flame. Otherwise the slide may break or you may damage the crystals.
3. Using a low power microscope or a good hand magnifier, study the form of the crystals. Make a sketch of each.
4. Mix 3 cc 0.1M potassium chromate solution, K_2CrO_4 , and 3 cc 0.1M barium chloride solution, $BaCl_2$, in a small test tube. Note results. Then filter the mixture, catching the filtrate in a clean test tube. Does filtering separate the precipitate from the liquid?
5. Heat the filtrate from step 4 to boiling. Boil 3-4 minutes. Filter again.

Part II: Molecular Crystals

1. When roll sulfur is dissolved in carbon disulfide, CS_2 , and the liquid is evaporated, rhombic crystals of sulfur form. Observe crystals prepared in this way with a magnifier or if possible with a microscope. Sketch one.
2. Sulfur also has another crystalline form. Place about 2 cc of powdered sulfur in a test tube. Fold a filter paper and put it in a funnel. Support the funnel in a second test tube. Heat the sulfur and the entire first test tube very gently until the sulfur is melted. (The upper portion of the tube must be warm so the sulfur does not solidify as you pour it.) Do *not* allow it to become hot enough to darken. It should be a pale yellow liquid. Pour this liquid into the filter paper. Blow across the surface. When a thin film has formed, open the filter paper and observe the needle-like crystals. These are prismatic, or monoclinic, crystals. Observe them with a magnifier. Sketch a sample.
3. Place a 50 cc beaker on an asbestos mat above a burner. Half-fill an evaporating dish with cold water. Put a small amount of iodine in the beaker. Then place the evaporating dish on the beaker and heat very gently. (Caution: Do *not* handle iodine with bare fingers.) Little heat is needed and the burner should be removed as soon as the iodine vaporizes. When you can see crystals of iodine on the bottom of the evaporating dish, remove it and pour out the water. Be careful not to damage the crystals. Examine these with a magnifier. Sketch a sample.



Part III: Metallic Crystals

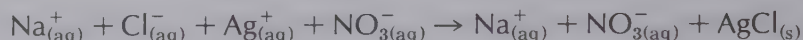
1. Place a 2 cm piece of copper wire on a glass slide. Cover it with several drops of silver nitrate solution. Watch as silver crystals form. Observe them with a magnifier and if possible with a microscope as they form. Sketch a sample.

QUESTIONS:

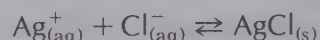
1. How could you use the process of crystallization to separate a mixture of two ionic salts?
2. Why is iodine a solid at room temperature, while the other halogens are liquid or gaseous? Use your knowledge of the Periodic Table and the trends represented.
3. Could you produce copper crystals by placing silver wire in a copper sulfate solution? Explain.
4. Propose a hypothesis to explain why some precipitates pass through filter paper.
5. What effect does heating a mixture containing a precipitate have upon the size of the crystals in the precipitate? Cite evidence for your answer.
6. What kind of properties are associated with the crystals of each bond type?

10-3 • IONIC REACTIONS IN WATER SOLUTIONS

Ionic solids dissolve in water to form ions. For example, sodium chloride, NaCl, releases Na^+ and Cl^- ions when dissolved in water. If another ionic solution is added to the sodium chloride solution, what happens? It depends upon the ions present in the second solution. If the solution is silver nitrate, containing Ag^+ and NO_3^- ions, a precipitate forms. We represent the reaction with an ionic equation:



Remember that in an ionic equation, the ions that are in a water solution have the subscript (aq), from the latin word *aqua* (water). The precipitate has a subscript (s), for solid. Silver chloride is relatively insoluble in water. Most of it settles out in solid form as a precipitate. This removes most of the Ag^+ ions and the Cl^- ions. The net ionic equation shows only the ions that combine to form a precipitate. All other ions are called spectator ions: they are unchanged. The net result of this reaction is an equilibrium described by the net ionic equation:



On the other hand, if we add a solution of potassium nitrate to the sodium chloride solution, the results are different:



In this case, no precipitate forms; all of the ions remain in solution. There is no net change.

How can we predict whether a precipitate will form when two solutions are mixed? There is no simple way. Many factors affect the solubility of solids. However, there are general guidelines, or rules, which are very useful in predicting whether a precipitate will form. In this experiment you will mix combinations of ionic solutions in many ways, and on the basis of your observations, formulate a set of rules for solubility. Though not infallible, these rules will be useful. Particularly, they will be important when you begin qualitative analysis.

PURPOSE:

To learn how ions react in aqueous solutions and to identify a set of "solubility patterns" which can be expressed as rules for use in predicting reactions. An additional purpose of this experiment is to provide practice in writing ionic equations.

PROCEDURE:

You should prepare a data table like the one shown opposite for all three sets of solutions. Decide which tests need to be performed and which would be a duplication of effort. Avoid wasting time on unnecessary tests.



Solutions, Set 1

BaCl_2
 Na_2SO_4
 $\text{Ba}(\text{NO}_3)_2$
 AgNO_3
 K_2CrO_4
 $\text{Pb}(\text{NO}_3)_2$

Solutions, Set 2

$(\text{NH}_4)_2\text{S}$
 CoCl_2
 $\text{Ni}(\text{NO}_3)_2$
 $\text{Cu}(\text{NO}_3)_2$
 FeCl_3
 NH_4OH

Solutions, Set 3

$\text{Ba}(\text{NO}_3)_2$
 CuCl_2
 AlCl_3
 AgNO_3
 NiCl_2
 Na_2CO_3

Sample Data Table

	$\text{Pb}(\text{NO}_3)_2$	K_2CrO_4	AgNO_3	$\text{Ba}(\text{NO}_3)_2$	Na_2SO_4	BaCl_2
BaCl_2						
Na_2SO_4						
$\text{Ba}(\text{NO}_3)_2$						
AgNO_3						
K_2CrO_4						
$\text{Pb}(\text{NO}_3)_2$						

1. Begin with any one of the three sets of solutions.
2. If you do not have spot plates like those in the photo, use a grease pencil or crayon to draw a grid on a piece of clear plastic that corresponds to your data tables. Make the squares at least 2×2 cm. Place the plastic over a sheet of black paper. Also have a sheet of white paper on hand to slide under the plastic. Some precipitates show up better against black; others, against white.
3. Place 1 drop of each of the two solutions being tested in a given square. (Note: Do *not* contaminate the solutions! Be careful not to touch the tip of a dropper to the liquid already on the plate.)
4. Observe carefully to see whether a precipitate forms. It sometimes takes a few seconds for a reaction to occur. Note the color of the precipitate in the proper section of your data table. If no precipitate forms, draw a straight line in the center of the square.
5. When you have finished observations on *one* set of solutions, write properly balanced chemical equations for all reactions that produced a precipitate. Write the color of the precipitate in the right margin after the equation. Try to determine which product is the precipitate. Use the following table of ions for writing equations:

Name of ion	Formula of ion, with charge
Aluminum	Al^{3+}
Ammonium	NH_4^+
Barium	Ba^{2+}
Cobalt	Co^{2+}
Copper	Cu^{2+}
Iron	Fe^{3+}
Lead	Pb^{2+}
Nickel	Ni^{2+}
Potassium	K^+
Silver	Ag^+
Sodium	Na^+
Carbonate	CO_3^{2-}
Chloride	Cl^-
Chromate	CrO_4^{2-}
Hydroxide	OH^-
Nitrate	NO_3^-
Sulfate	SO_4^{2-}
Sulfide	S^{2-}

6. Have your equations checked by your instructor *before* going on to another set.
7. After your equations have been approved, repeat the above steps for the other two sets of solutions.

ANALYSIS OF DATA:

1. After equations for precipitation reactions have been written and approved for all three sets of solutions, examine your data. It might be well to tabulate the number of times precipitates contained Ba^{2+} , for example. A tabulation chart can be very helpful in determining the ions most likely to form precipitates.
2. After you have examined your data, formulate a set of solubility rules based on the patterns you have observed. These will serve as guidelines in future experiments. For example, if you find Ag^+ and Cl^- tend to form a precipitate whenever in the same solution, you might state as a general rule: Whenever silver ion and chloride ion come together in a solution, a precipitate will form.
3. What generalizations can you make about the color that certain ions impart to solutions and to precipitates? This will be helpful when you begin your study of qualitative analysis.

11-1 • IDENTIFICATION OF CERTAIN ANIONS

In this simplified procedure for qualitative analysis the anions are divided into four groups. In this experiment you will perform tests for the common anions in which you will get positive results. You will then know when these ions are present when you undertake the analysis of unknowns.

PURPOSE:

To observe positive tests for various anions and thus, know how to recognize the presence of these anions in unknowns.

PROCEDURE AND RESULTS:

Maintain a very careful record of what you observe in the following tests. You will want to refer to it when you work with unknowns. Try to devise a table that summarizes all the observations.

Part A: Tests for Group I Anions (CO_3^{2-} , NO_2^- , SO_3^{2-} , S^{2-})

General Test for Group I Anions

1. Place 2 cc of each of the following solutions in separate test tubes: sodium carbonate, sodium nitrite, sodium sulfite, and sodium sulfide.
2. Add 5 drops HCl to each solution. Effervescence indicates the presence of Group I anions.

Test for Carbonate

1. Place 2 cc sodium carbonate in a test tube. Add a few drops of AgNO_3 solution. Record result.

Tests for Nitrite

1. Place 2 cc sodium nitrite in a test tube. Add a few drops AgNO_3 solution. Record result.
2. Place 2 cc sodium nitrite in a test tube. Add 10 drops H_2SO_4 (6M). Record result.

Test for Sulfite

1. Place 2 cc sodium sulfite in a test tube. Add 3 drops concentrated H_2SO_4 . Cautiously note odor.

Test for Sulfide

1. Pour 2 cc sodium sulfide in a test tube. Add 2 cc HCl. Note odor. Record result.
2. Pour 2 cc sodium sulfide in a test tube. Add a few drops AgNO_3 . Note result.

Part B: Tests for Group II Anions (CrO_4^{2-} , PO_4^{3-} , SO_4^{2-})

General Test for Group II Anions

1. Place 2 ml of each of the following solutions in separate test tubes: potassium chromate, sodium phosphate, and sodium sulfate.
2. To each test tube add 10 drops BaCl_2 solution. Record result.

Test for Chromate

1. Place 2 ml potassium chromate in a test tube. Add 5 drops AgNO_3 . Record result.

Tests for Phosphate

1. Place 2 ml sodium phosphate in each of two separate test tubes. To one, add 10 drops BaCl_2 solution. White precipitate indicates phosphate. Add 1 ml HCl to mixture. If white precipitate dissolves, phosphate is confirmed.
2. To the second test tube, add 10 drops AgNO_3 . A yellow precipitate indicates phosphate. Add 1 ml HNO_3 . If precipitate dissolves, phosphate is confirmed.

Test for Sulfate

1. Place 2 ml sulfate in a test tube. Add 10 drops BaCl_2 . A white precipitate indicates sulfate. Add 1 ml HCl . If precipitate does *not* dissolve, sulfate is confirmed.

Part C: Tests for Group III Anions (Cl^- , Br^- , I^-)

General Test for Group III Anions

1. Place 2 ml of each of the following solutions in separate test tubes: sodium chloride, sodium bromide, sodium iodide.
2. To each solution, add 10 drops AgNO_3 . Note precipitate color very carefully.

Test for Chloride

1. To the chloride precipitate from above, add 2 ml NH_4OH . If precipitate dissolves, chloride is indicated. Add 2 ml HNO_3 , or until solution is acidic to litmus paper. Reappearance of white precipitate confirms the presence of Cl^- ion.

Test for Bromide and Iodide

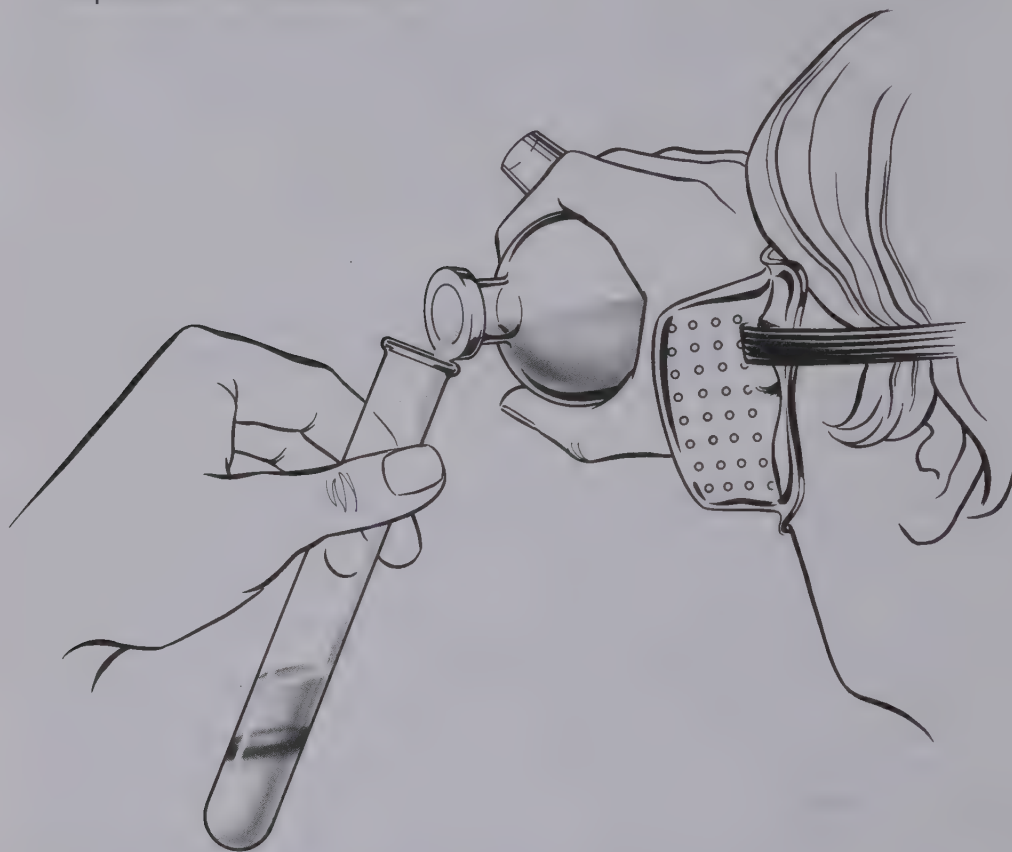
1. Place 2 ml sodium bromide in one test tube and 2 ml sodium iodide in a second test tube. To each solution, add 1 ml chlorine water and 1 ml CCl_4 . Shake well. Orange layer in the bottom indicates bromide. Violet color indicates iodide. Compare with color plate found on page 122 of text.

Part D: Tests for Group IV Anions (NO_3^- , $\text{C}_2\text{H}_3\text{O}_2^-$, ClO_3^-)

1. Place 2 ml of each of the following solutions in separate test tubes: sodium nitrate, sodium acetate, sodium chlorate.

Test for Nitrate

1. Wear goggles while performing this test. To the test tube containing sodium nitrate, add 2 ml freshly prepared FeSO_4 solution. Mix well. Hold the test tube at a 45° angle and carefully trickle concentrated H_2SO_4 slowly down the inside of the tube. About 1 ml is sufficient. When the last of the acid has reached the bottom, carefully move the tube to an upright position. The appearance of a brown ring between the two layers confirms the presence of the nitrate ion.

**Test for Chlorate**

1. To the test tube containing sodium chlorate, add 10 drops concentrated H_2SO_4 . Heat gently. Cautiously note the color of the escaping gas. Green gas confirms chlorate.

Test for Acetate

1. To the test tube containing sodium acetate, add 10 drops H_2SO_4 . Heat gently. Cautiously note odor. Odor of vinegar confirms the presence of the acetate ion.

QUESTIONS:

1. By what odor can you identify the sulfite ion? The acetate ion?
2. By what color of gas released can you identify the chlorate ion? The nitrite ion?
3. By what color in the CCl_4 layer can you identify the bromide ion? The iodide ion?
4. Explain the function of the chlorine water and the carbon tetrachloride reagents in the bromide-iodide tests.
5. Both phosphates and sulfates form white precipitates upon addition of barium chloride solution. What specific test determines which ion is present?
6. Chlorides, bromides, and iodides all form precipitates upon the addition of AgNO_3 . What precipitate color helps narrow the possibilities for the anions present?

EQUATIONS:

Write complete ionic equations for the following reactions.

1. $\text{Na}_2\text{CO}_3 + \text{HCl}$
2. $\text{NaC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{SO}_4$
3. $\text{NaNO}_2 + \text{H}_2\text{SO}_4$
4. $\text{Na}_2\text{S} + \text{HCl}$
5. $\text{Na}_2\text{SO}_3 + \text{H}_2\text{SO}_4$
6. $\text{Na}_2\text{SO}_4 + \text{BaCl}_2$
7. $\text{Na}_3\text{PO}_4 + \text{BaCl}_2$
8. $\text{K}_2\text{CrO}_4 + \text{BaCl}_2$
9. $\text{NaCl} + \text{AgNO}_3$

11-2. ELEMENTARY QUALITATIVE ANALYSIS

You have studied the theory of qualitative analysis and have been acquainted with the tests for some of the cations and anions. In this experiment you will be given a number of unknowns to identify. Use the concepts and principles you have learned. Follow instructions closely but not blindly. Keep careful records of all tests and the results.

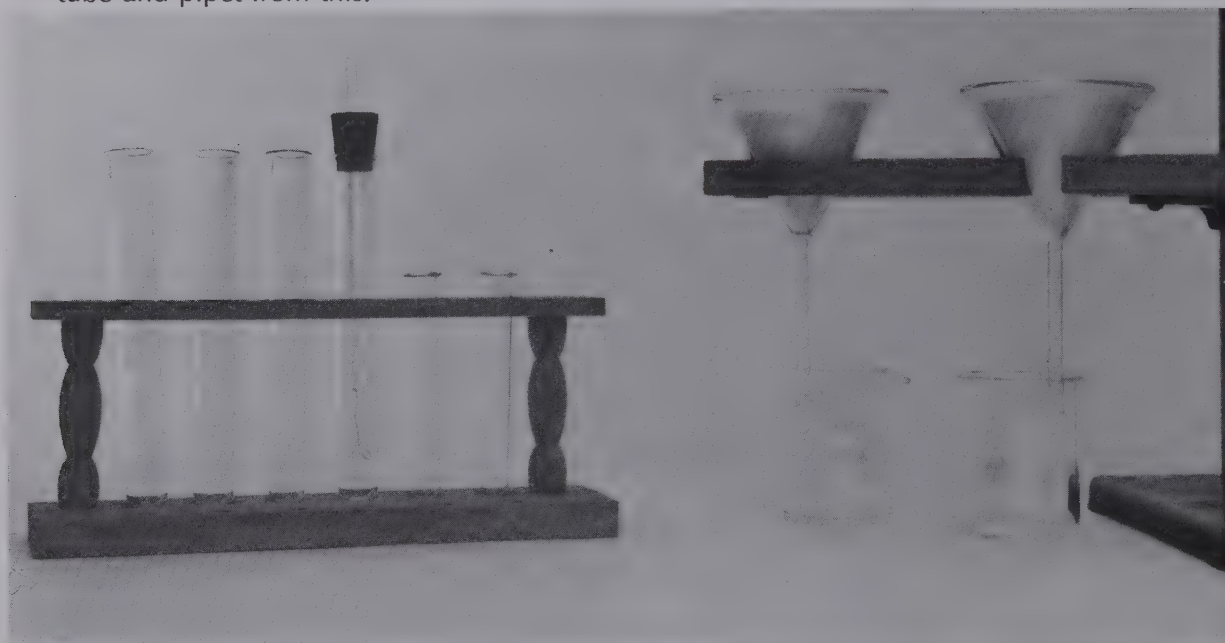
PURPOSE:

To identify the ions in aqueous solutions by following a systematic procedure of analysis and in so doing to apply chemical theory to interpret the reactions that occur.

GENERAL RULES OF PROCEDURE:

For your convenience, the general rules of procedure are listed. Study them carefully; refer to them at intervals.

1. Unless otherwise directed, use 1 cc portions of the unknown solution in 13×100 mm test tubes.
2. Unless otherwise directed, add about 10 drops of the test reagent to 1 cc of the unknown.
3. Cleanliness is of utmost importance. Keep a beaker of soapy water and a test tube brush on your desk at all times. As soon as a test tube has been emptied, wash it, rinse it well with tap water, and turn it upside down to drain. In this way you will always have clean equipment with which to work. Should it be necessary to use the test tube immediately, rinse it with distilled water and shake to remove excess water.
4. Never insert a stirring rod or pipet into a stock bottle of reagent. This can cause contamination. Pour a small volume of the reagent into a small test tube and pipet from this.



5. Always label test tubes and beakers so there is no doubt about their contents. Use labels that are applied by pressure. These stick to dry glassware, but do not have to be scraped off.
6. Keep careful, dated records of the code number of each unknown, the group and confirmatory tests performed on it with results, all problems encountered, etc. It is especially important to note test results that seem questionable. These can serve as clues for a recheck if it becomes desirable. The notebook used for qualitative analysis records should be permanently bound with no loose sheets.

If it seems that you have recorded incorrect data, draw a line through the entry, but leave it legible. Frequently it is helpful to refer to data that has been discarded prematurely.

7. Write balanced chemical equations for all reactions. Underline the formula for the precipitate and note its color to the right of the equation.
8. Remember that hydrogen sulfide, H_2S , is poisonous! Do *not* use it carelessly.
9. To test the acidity or alkalinity of a solution, dip a clean stirring rod into the solution and touch it to a piece of litmus paper placed on a clean watch glass. The same piece of litmus can then be used for many such tests. This also conserves solution, as about 1 ml of liquid is needed to saturate one piece of litmus paper.
10. After using a platinum wire, clean it and return it to the main reagent table for others to use. These wires are expensive and are not usually provided in unlimited quantities.
11. In cation analysis, *all* confirmatory tests must be run on the original unknown solution.

REPORTING RESULTS

Follow the rules for reporting your work on an unknown. Time will be saved and misunderstandings will be minimized. Your instructor will want to know your results and your reasons for reaching your conclusions.

1. Whenever possible, you will need a positive group test and at least two positive confirmatory tests before reporting the identification of cations. In many cases, only one test is required for identification of anions. *All positive tests must be shown to the instructor when reporting.*
2. Have your cation confirmed by the instructor before beginning tests on the anion. This is important. In most cases you will precipitate the metallic ions and discard them to prevent interference with anion tests. It is wise to have the cations confirmed before it is too late for a recheck.
3. When reporting your conclusions to the instructor:
 - a. Report the names of the ions for which you have positive identification.
 - b. Tell the reagent used for the group test and the name of the precipitate.
 - c. Tell the reagents used for confirmatory tests and the names of all precipitates.

- d. If flame tests, bead tests, or cobalt nitrate tests are used, show these results to the instructor.
 - e. Be prepared to explain the function of each reagent used. If uncertain, try to find out by consulting your textbook or other references. If this fails, ask your instructor. Independence and initiative are valuable assets.
4. If your instructor requires a written report, follow the form requested.

PRELIMINARY OBSERVATIONS ON AN UNKNOWN

1. Test the pH with hydron paper or the general acidity with litmus paper.
2. Check the color of the solid salt by evaporating one drop of solution on a glass slide. (*Caution:* Do not overheat; the slide may break and the residue may discolor.)

TESTING FOR GROUP I CATIONS

(Ag^+ , Hg_2^{2+} , Pb^{2+})

Group I ions are precipitated by 3M HCl. *Note:* As you perform this first test, you will obtain evidence about the presence of Group I anions. If effervescence occurs when HCl is added to the unknown, this indicates the presence of Group I anions. Be sure to note this in your record.

1. Pour 1 ml of unknown into a test tube. Add 10 drops 3M HCl. If no precipitate forms, use this solution for the Group II test.
2. If a white precipitate forms, the presence of Group I cations is indicated. Assume the presence of all ions and proceed as follows.
 - a. Add 5 ml of unknown to the test tube containing group precipitates. Add 3 ml HCl. This completes precipitation of sufficient material to permit separation of the ions.
 - b. Filter mixture by setting funnel in clean test tube. Wash filter paper and precipitates with 5 ml distilled water. Discard filtrate.
 - c. Place funnel containing precipitates in clean test tube. Pour 5 ml hot, distilled water through the filter paper. Remove the funnel to a clean test tube. Add 1 ml K_2CrO_4 to the filtrate. A yellow precipitate confirms Pb^{2+} .
 - d. If Pb^{2+} is present, repeat step c until no more yellow precipitate forms. Discard additional lead precipitates.
 - e. Pour 5 ml 6M $\text{NH}_{3(\text{aq})}$ through the filter paper, catching the filtrate in a clean test tube. If the precipitate turns black, the presence of Hg_2^{2+} is confirmed.
 - f. To the filtrate from step e, add 6M HNO_3 until the solution is acidic to litmus. This usually takes about 5 ml. A white precipitate confirms the presence of Ag^+ .
 - g. In this simplified procedure, you will not have cations from more than one group in one unknown. Report your results to the instructor for confirmation before proceeding with the tests for anions.

TESTING FOR GROUP II CATIONS (Cu^{2+} , Hg^{2+} , Bi^{2+} , Sb^{3+})

Group II ions are precipitated by H_2S in 3 M HCl solution. Bubble H_2S into the acidified solution from step 1, Group I. If a precipitate forms, the presence of Group II is indicated. Refer to the following instructions and do confirmatory tests on the original unknown for the ion indicated by the color of the precipitate. If no precipitate forms, discard solution and proceed to test for Group III.

Confirmatory tests - Group II

Orange precipitate suggests Antimony

- If the H_2S precipitate is soluble in $(\text{NH}_4)_2\text{S}$, Antimony is confirmed.

Black precipitate suggests Copper (II)

- Add 3 drops $\text{NH}_{3(\text{aq})}$. A blue-white precipitate that dissolves in excess $\text{NH}_{3(\text{aq})}$ to make deep blue solution confirms Copper II.
- Green flame test also confirms Copper (II).

Yellow-black precipitate suggests Mercury (II)

- Add 3 drops KI. If red precipitate forms, then dissolves in excess KI, Mercury (II) is confirmed.

Brown precipitate suggests Bismuth

- Add 3 drops KI. A brown precipitate suggests Bismuth.
- Add 3 drops K_2CrO_4 . A yellow precipitate confirms Bismuth.

TESTING FOR GROUP III CATIONS (Cr^{3+} , Al^{3+} , Fe^{3+})

Group III ions are precipitated by $\text{NH}_{3(\text{aq})}$ in the presence of NH_4Cl . Pour 2 ml unknown into a test tube. Add 10 drops NH_4Cl . Then add 10 drops $\text{NH}_{3(\text{aq})}$. If a precipitate forms, the presence of Group III is indicated. Refer to the following instructions and do confirmatory tests on the original unknown for the ion indicated by the color of the precipitate. If no precipitate forms, discard solution and proceed to tests for Group IV.

Confirmatory tests - Group III

Gray-green precipitate suggests Chromium

- Add $(\text{NH}_4)_2\text{CO}_3$. A greenish precipitate suggests Chromium.
- A green borax bead confirms Chromium.

White, gelatinous precipitate suggests Aluminum

- Add $(\text{NH}_4)_2\text{CO}_3$. If white, gelatinous precipitate persists, it is suggestive of Aluminum.
- Blue cobalt nitrate test confirms Aluminum.

Red-brown precipitate suggests Iron (III)

- Add KSCN. A red coloration suggests Iron III.
- Add $K_4Fe(CN)_6$. A deep blue precipitate confirms Iron III.

TESTING FOR GROUP IV CATIONS
(Co^{2+} , Ni^{2+} , Zn^{2+} , Mn^{2+})

Group IV ions are precipitated by $(NH_4)_2S$ in the presence of NH_4Cl . Pour 2 ml unknown into a test tube. Add 10 drops NH_4Cl . Then add 10 drops $(NH_4)_2S$. If precipitate forms, the presence of Group IV is indicated. Refer to the following instructions and run confirmatory tests on the original unknown for the ion indicated by the color of the precipitate. If no precipitate forms, discard solution and proceed to test for Group V.

Confirmatory tests - Group IV**Black precipitate suggests Cobalt**

- Add 4 drops $NH_{3(aq)}$. A green precipitate dissolving in excess $NH_{3(aq)}$ suggests Cobalt.
- A deep blue borax bead confirms Cobalt.

White precipitate suggests Zinc

- Add 4 drops $NH_{3(aq)}$. A white precipitate, soluble in acetic acid, suggests Zinc.
- Green cobalt nitrate test confirms Zinc.

Black precipitate also suggests Nickel

- Add 4 drops $NH_{3(aq)}$. A blue precipitate dissolving in excess $NH_{3(aq)}$ suggests Nickel.
- A brown borax bead confirms Nickel.

Salmon (pinkish) precipitate suggests Manganese

- Add 4 drops $NH_{3(aq)}$. A white precipitate rapidly oxidizing to brown suggests Manganese.
- Lilac borax bead confirms Manganese.

TESTING FOR GROUP V CATIONS
(Sr^{2+} , Ba^{2+} , Ca^{2+})

Group V ions are precipitated by $(NH_4)_2CO_3$ in the presence of NH_4Cl . Pour 2 ml unknown into a test tube. Add 10 drops NH_4Cl . Add 10 drops $(NH_4)_2CO_3$. If white precipitate forms, the presence of Group V is indicated. Run confirmatory tests as follows.

Confirmatory Tests - Group V

White precipitate suggests Strontium

- a. Bright red flame test confirms Strontium.

White precipitate also suggests Barium

- a. Greenish-yellow flame test further suggests Barium.
- b. Add 5 drops $\text{HC}_2\text{H}_3\text{O}_2$ and 5 drops K_2CrO_4 . A yellow precipitate confirms Barium.

White precipitate also suggests Calcium

- a. Orange flame test further suggests Calcium.
- b. Add 5 drops $(\text{NH}_4)_2\text{C}_2\text{O}_4$; a white precipitate confirms Calcium.

TESTING FOR GROUP VI CATIONS (K^+ , Na^+ , NH_4^+)

Group VI ions are not precipitated, so special tests must be used to detect them.

Potassium

- a. Violet flame test indicates Potassium.

Sodium

- a. Persistent yellow flame indicates Sodium.

Ammonium

- a. To 2 ml unknown, add 10 drops NaOH . Warm; typical odor of ammonia confirms presence.

TESTING FOR ANIONS

Note: If effervescence occurred when you added HCl to test for Group I cations, the presence of Group I anions is indicated. The original unknown solution is used for these tests. Proceed directly to run confirmatory tests for this Group I. If effervescence did *not* occur, proceed as directed below.

A. Preparation of unknown solution for anion tests

1. If cations in Groups I through V are present, these *must* be removed from the solution before testing for anions. They will interfere with the tests. To conserve time, remove the metallic ions from 10–15 ml of the unknown at once. The resulting solution will be tested in small portions as needed. If Group VI cations are present, you may go on to test for anions without removing the metallic ions.

2. *To remove metallic ions:* Pour 10–15 ml unknown solution into a 250 ml beaker. Add Na_2CO_3 until precipitation appears to stop (this may require 10 ml of reagent). Mix thoroughly. Heat to boiling and filter, catching filtrate in large, clean test tube. Discard residue on filter paper. To be sure all the metallic ions have been removed, add a few more drops of Na_2CO_3 . If more precipitate forms, repeat above procedure. If no precipitate forms, label filtrate and use as needed to test for the Group II, III, and IV anions.

B. Testing for groups

1. Use the directions described in Experiment 11-1 to test for the Group II, III, and IV anions.

APPENDICES

VAPOR PRESSURE OF WATER

Vapor Pressure of Water, in Torr, at Various Temperatures	
Temperature, °C	Vapor Pressure, Torr (to the nearest whole number)
0	4
5	7
10	9
15	13
17	14
18	15
19	16
20	18
21	19
22	20
23	21
24	22
25	24
30	32
35	42
40	55
50	92
60	149
70	234
80	355
90	526
100	760

CHARGES AND OXIDATION NUMBERS OF COMMON IONS

Oxidation Number	Name of Ion	Formula	Oxidation Number	Name of Ion	Formula
+1	ammonium	NH_4^+	-1	acetate	$\text{C}_2\text{H}_3\text{O}_2^-$
	copper(I)	Cu^+		bromide	Br^-
	hydrogen	H^+		chlorate	ClO_3^-
	lithium	Li^+		chloride	Cl^-
	potassium	K^+		chlorite	ClO_2^-
	silver	Ag^+		fluoride	F^-
	sodium	Na^+		hydrogen carbonate	HCO_3^-
				hydroxide	OH^-
+2	barium	Ba^{2+}		hypochlorite	ClO^-
	calcium	Ca^{2+}		iodide	I^-
	chromium(II)	Cr^{2+}		nitrate	NO_3^-
	cobalt	Co^{2+}		nitrite	NO_2^-
	copper(II)	Cu^{2+}		perchlorate	ClO_4^-
	iron(II)	Fe^{2+}		permanganate	MnO_4^-
	lead(II)	Pb^{2+}		thiocyanate	SCN^-
	magnesium	Mg^{2+}	-2	carbonate	CO_3^{2-}
	manganese(II)	Mn^{2+}		chromate	CrO_4^{2-}
	mercury(I)	Hg_2^{2+}		dichromate	$\text{Cr}_2\text{O}_7^{2-}$
	mercury(II)	Hg^{2+}		oxalate	$\text{C}_2\text{O}_4^{2-}$
	tin(II)	Sn^{2+}		sulfate	SO_4^{2-}
	zinc	Zn^{2+}		sulfite	SO_3^{2-}
				thiosulfate	$\text{S}_2\text{O}_3^{2-}$
+3	aluminum	Al^{3+}	-3	ferricyanide	$\text{Fe}(\text{CN})_6^{3-}$
	antimony(III)	Sb^{3+}		phosphate	PO_4^{3-}
	chromium(III)	Cr^{3+}	-4	ferrocyanide	$\text{Fe}(\text{CN})_6^{4-}$
	iron(III)	Fe^{3+}			
+4	tin(IV)	Sn^{4+}			

Periodic Table of the Elements

1 H Hydrogen 1.008

IA	IIA
3 Li Lithium 6.939	4 Be Beryllium 9.012
11 Na Sodium 22.99	12 Mg Magnesium 24.31

22	—	Atomic number
	Ti	Symbol of element
	Titanium	Name of element
	47.90	Atomic weight (based on ^{12}C scale)

IIIB	IVB	VB	VIB	VIIB	VIIIB	VIIIB	VIIIB	VIIIB
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.90	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.94	43 Tc Technetium (99)	44 Ru Ruthenium 101.1	45 Rh Rhodium 102.9
55 Cs Cesium 132.9	56 Ba Barium 137.3		72 Hf Hafnium 178.5	73 Ta Tantalum 180.9	74 W Wolfram 183.9	75 Re Rhenium 186.2	76 Os Osmium 190.2	77 Ir Iridium 192.2
87 Fr Francium (223)	88 Ra Radium (226)							
Lanthanide series			57 La Lanthanum 138.9	58 Ce Cerium 140.1	59 Pr Praseodymium 140.9	60 Nd Neodymium 144.2	61 Pm Promethium (145)	62 Sm Samarium 150.4
Actinide series			89 Ac Actinium (227)	90 Th Thorium 232.0	91 Pa Protactinium (231)	92 U Uranium 238.0	93 Np Neptunium (237)	94 Pu Plutonium (242)

										VIIIA
										2 He Helium 4.003
										IIIA IVA VA VIA VIIA
										5 B Boron 10.81
										6 C Carbon 12.01
										7 N Nitrogen 14.00
										8 O Oxygen 15.9994
										9 F Fluorine 18.998
										10 Ne Neon 20.183
										13 Al Aluminum 26.98
										14 Si Silicon 28.09
										15 P Phosphorus 30.97
										16 S Sulfur 32.06
										17 Cl Chlorine 35.45
										18 Ar Argon 39.95
										IB IIB
28 Ni Nickel 58.71	29 Cu Copper 63.54	30 Zn Zinc 65.37	31 Ga Gallium 69.72	32 Ge Germanium 72.59	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.91	36 Kr Krypton 83.80		
46 Pd Palladium 106.4	47 Ag Silver 107.9	48 Cd Cadmium 112.4	49 In Indium 114.8	50 Sn Tin 118.7	51 Sb Antimony 121.8	52 Te Tellurium 127.6	53 I Iodine 126.9	54 Xe Xenon 131.3		
78 Pt Platinum 195.1	79 Au Gold 197.0	80 Hg Mercury 200.6	81 Tl Thallium 204.4	82 Pb Lead 207.2	83 Bi Bismuth 209.0	84 Po Polonium (210)	85 At Astatine (210)	86 Rn Radon (222)		

63 Eu Europium 152.0	64 Gd Gadolinium 157.3	65 Tb Terbium 158.9	66 Dy Dysprosium 162.5	67 Ho Holmium 164.9	68 Er Erbium 167.3	69 Tm Thulium 168.9	70 Yb Ytterbium 173.0	71 Lu Lutetium 175.0
-------------------------------	---------------------------------	------------------------------	---------------------------------	------------------------------	-----------------------------	------------------------------	--------------------------------	-------------------------------

95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (249)	98 Cf Californium (251)	99 Es Einsteinium (254)	100 Fm Fermium (252)	101 Md Mendelevium (256)	102 No Nobelium (254)	103 Lw Lawrencium (256)
--------------------------------	-----------------------------	--------------------------------	----------------------------------	----------------------------------	-------------------------------	-----------------------------------	--------------------------------	----------------------------------

Electronegativity Values Represent an Arbitrary Scale
to Indicate the Attraction of an Atom for Electrons
Involved in the Formation of a Chemical Bond.

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36	37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54	55	56	57 *	58	59	60	61	62	63	64	65	66	67	68	69	70	71	72	73	74	75	76	77	78	79	80	81	82	83	84	85	86	87	88	89 **	90	91	92	93	94	95	96	97	98	99	100	101	102	103	104	105	106	107	108	109	110	111	112	113	114	115	116	117	118	119	120	121	122	123	124	125	126	127	128	129	130	131	132	133	134	135	136	137	138	139	140	141	142	143	144	145	146	147	148	149	150	151	152	153	154	155	156	157	158	159	160	161	162	163	164	165	166	167	168	169	170	171	172	173	174	175	176	177	178	179	180	181	182	183	184	185	186	187	188	189	190	191	192	193	194	195	196	197	198	199	200	201	202	203	204	205	206	207	208	209	210	211	212	213	214	215	216	217	218	219	220	221	222	223	224	225	226	227	228	229	230	231	232	233	234	235	236	237	238	239	240	241	242	243	244	245	246	247	248	249	250	251	252	253	254	255	256	257	258	259	260	261	262	263	264	265	266	267	268	269	270	271	272	273	274	275	276	277	278	279	280	281	282	283	284	285	286	287	288	289	290	291	292	293	294	295	296	297	298	299	300	301	302	303	304	305	306	307	308	309	310	311	312	313	314	315	316	317	318	319	320	321	322	323	324	325	326	327	328	329	330	331	332	333	334	335	336	337	338	339	340	341	342	343	344	345	346	347	348	349	350	351	352	353	354	355	356	357	358	359	360	361	362	363	364	365	366	367	368	369	370	371	372	373	374	375	376	377	378	379	380	381	382	383	384	385	386	387	388	389	390	391	392	393	394	395	396	397	398	399	400	401	402	403	404	405	406	407	408	409	410	411	412	413	414	415	416	417	418	419	420	421	422	423	424	425	426	427	428	429	430	431	432	433	434	435	436	437	438	439	440	441	442	443	444	445	446	447	448	449	450	451	452	453	454	455	456	457	458	459	460	461	462	463	464	465	466	467	468	469	470	471	472	473	474	475	476	477	478	479	480	481	482	483	484	485	486	487	488	489	490	491	492	493	494	495	496	497	498	499	500	501	502	503	504	505	506	507	508	509	510	511	512	513	514	515	516	517	518	519	520	521	522	523	524	525	526	527	528	529	530	531	532	533	534	535	536	537	538	539	540	541	542	543	544	545	546	547	548	549	550	551	552	553	554	555	556	557	558	559	560	561	562	563	564	565	566	567	568	569	570	571	572	573	574	575	576	577	578	579	580	581	582	583	584	585	586	587	588	589	590	591	592	593	594	595	596	597	598	599	600	601	602	603	604	605	606	607	608	609	610	611	612	613	614	615	616	617	618	619	620	621	622	623	624	625	626	627	628	629	630	631	632	633	634	635	636	637	638	639	640	641	642	643	644	645	646	647	648	649	650	651	652	653	654	655	656	657	658	659	660	661	662	663	664	665	666	667	668	669	670	671	672	673	674	675	676	677	678	679	680	681	682	683	684	685	686	687	688	689	690	691	692	693	694	695	696	697	698	699	700	701	702	703	704	705	706	707	708	709	710	711	712	713	714	715	716	717	718	719	720	721	722	723	724	725	726	727	728	729	730	731	732	733	734	735	736	737	738	739	740	741	742	743	744	745	746	747	748	749	750	751	752	753	754	755	756	757	758	759	760	761	762	763	764	765	766	767	768	769	770	771	772	773	774	775	776	777	778	779	780	781	782	783	784	785	786	787	788	789	790	791	792	793	794	795	796	797	798	799	800	801	802	803	804	805	806	807	808	809	810	811	812	813	814	815	816	817	818	819	820	821	822	823	824	825	826	827	828	829	830	831	832	833	834	835	836	837	838	839	840	841	842	843	844	845	846	847	848	849	850	851	852	853	854	855	856	857	858	859	860	861	862	863	864	865	866	867	868	869	870	871	872	873	874	875	876	877	878	879	880	881	882	883	884	885	886	887	888	889	890	891	892	893	894	895	896	897	898	899	900	901	902	903	904	905	906	907	908	909	910	911	912	913	914	915	916	917	918	919	920	921	922	923	924	925	926	927	928	929	930	931	932	933	934	935	936	937	938	939	940	941	942	943	944	945	946	947	948	949	950	951	952	953	954	955	956	957	958	959	960	961	962	963	964	965	966	967	968	969	970	971	972	973	974	975	976	977	978	979	980	981	982	983	984	985	986	987	988	989	990	991	992	993	994	995	996	997	998	999	1000	1001	1002	1003	1004	1005	1006	1007	1008	1009	1010	1011	1012	1013	1014	1015	1016	1017	1018	1019	1020	1021	1022	1023	1024	1025	1026	1027	1028	1029	1030	1031	1032	1033	1034	1035	1036	1037	1038	1039	1040	1041	1042	1043	1044	1045	1046	1047	1048	1049	1050	1051	1052	1053	1054	1055	1056	1057	1058	1059	1060	1061	1062	1063	1064	1065	1066	1067	1068	1069	1070	1071	1072	1073	1074	1075	1076	1077	1078	1079	1080	1081	1082	1083	1084	1085	1086	1087	1088	1089	1090	1091	1092	1093	1094	1095	1096	1097	1098	1099	1100	1101	1102	1103	1104	1105	1106	1107	1108	1109	1110	1111	1112	1113	1114	1115	1116	1117	1118	1119	1120	1121	1122	1123	1124	1125	1126	1127	1128	1129	1130	1131	1132	1133	1134	1135	1136	1137	1138	1139	1140	1141	1142	1143	1144	1145	1146	1147	1148	1149	1150	1151	1152	1153	1154	1155	1156	1157	1158	1159	1160	1161	1162	1163	1164	1165	1166	1167	1168	1169	1170	1171	1172	1173	1174	1175	1176	1177	1178	1179	1180	1181	1182	1183	1184	1185	1186	1187	1188	1189	1190	1191	1192	1193	1194	1195	1196	1197	1198	1199	1200	1201	1202	1203	1204	1205	1206	1207	1208	1209	1210	1211	1212	1213	1214	1215	1216	1217	1218	1219	1220	1221	1222	1223	1224	1225	1226	1227	1228	1229	1230	1231	1232	1233	1234	1235	1236	1237	1238	1239	1240	1241	1242	1243	1244	1245	1246	1247	1248	1249	1250	1251	1252	1253	1254	1255	1256	1257	1258	1259	1260	1261	1262	1263	1264	1265	1266	1267	1268	1269	1270	1271	1272	1273	1274	1275	1276	1277	1278	1279	1280	1281	1282	1283	1284	1285	1286	1287	1288	1289	1290	1291	1292	1293	1294	1295	1296	1297	1298	1299	1300	1301	1302	1303	1304	1305	1306	1307	1308	1309	1310	1311	1312	1313	1314	1315	1316	1317	1318	1319	1320	1321	1322	1323	1324	1325	1326	1327	1328	1329	1330	1331	1332	1333	1334	1335	1336	1337	1338	1339	1340	1341	1342	1343	1344	1345	1346	1347	1348	1349	1350	1351	1352	1353	1354	1355	1356	1357	1358	1359	1360	1361	1362	1363	1364	1365	1366	1367	1368	1369	1370	1371	1372	1373	1374	1375	1376	1377	1378	1379	1380	1381	1382	1383	1384	1385	1386	1387	1388	1389	1390	1391	1392	1393	1394	1395	1396	1397	1398	1399	1400	1401	1402	1403	1404	1405	1406	1407	1408	1409	1410	1411	1412	1413	1414	1415	1416	1417	1418	1419	1420	1421	1422	1423	1424	1425	1426	1427	1428	1429	1430	1431	1432	1433	1434	1435	1436	1437	1438	1439	1440	1441	1442	1443	1444	1445	1446	1447	1448	1449	1450	1451	1452	1453	1454	1455	1456	1457	1458	1459	1460	1461	1462	1463	1464	1465	1466	1467	1468	1469	1470	1471	1472	1473	1474	1475	1476	1477	1478	1479	1480	1481	1482	1483</
---	---	---	---	---	---	---	---	---	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	------	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	----	-------	----	----	----	----	----	----	----	----	----	----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	-----	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	------	--------

[illegible]

INDEX

(Page numbers in *italic* refer to the text *Keys to Chemistry*.)

acetate: test for, 124
 acid(s): 59–60; 74, 97, 101–106, 209–211, 292–293
 acid-base indicators, 61; 107, 113
 acidity, 128: 102, 110–115
 adsorption: selective, 71
 aluminum: tests for, 129
 ammonium: test for, 131
 anions, 122–125, 131–132; 273
 antimony: tests for, 129
 atom(s): number of particles in, 75
 size of, 75–76; 156
 atomic number: atomic radius and, 75–76; 164–168, 181–182
 ionization energy and, 77; 164–168, 179–180
 atomic radius: atomic number and, 75–76; 164–168, 181–182
 atomic weights: table of, v; 305

barium: tests for, 130–131
 bases, 59–60; 98, 106–108
 reaction with metal, 90
 bismuth: tests for, 129
 black-box experiments, 66; 6, 126–129
 boiling, 49–50; 66–71
 boiling temperature, 50; 67–71
 bond(s), 88; 177–196
 bond types, 179, 185–192, 195–196
 physical properties and, 86–87
 solution process and, 84–85
 borax bead test, 68–69; 277–278
 bromide: tests for, 123
 buffers and buffer solutions, 63–65; 112, 116, 275–276

calcium: tests for, 130–131
 calibration, 8, 10
 calorie: defined, 88; 16
 carbonate: test for, 122, 103
 carbon dioxide: tests for, 25
 carbon tetrachloride, 85
 cations, 80–82, 128–132; 273, 276–278
 centrifuge, 15
 chemical change, 3, 28–29, 40, 50–53
 evidences of, 39–40
 water promotes, 52; 74
 chemical equations: see equations

chemical properties, 39; 39
 chemical reactions: see reactions
 chemistry handbook, 35–36
 chlorate: test for, 124
 chloride: test for, 123
 chromate: test for, 123
 chromatography, 71–72
 chromium: test for, 129
 cobalt: test for, 130
 cobalt nitrate test, 69–70; 277–278
 color: in indicators, 61–62; 121
 ion identification by, 67–70; 122, 276–279
 as test, 122–125, 128–131
 common ion: effects of, 64; 115–116, 275
 hydrolysis, buffers, and, 63–65
 compound, 38, 43–44; 38, 46–47, 50
 concentration, 90, 91–93; 81, 104, 112
 copper: tests for, 129
 crystals, 39, 73, 248, 250, 255–257
 types of, 115–117; 256–260
 water in, 113–114

data: tables of, 2, 28
 decomposition reactions, 95; 218–219

decrepitation, 113–114
 deliquescence, 113; 249
 density, 31–32; 24–27
 diameter: circumference and, 30
 drying residues, 15
 dynamic equilibrium, 47–48; 65, 100

effervescence, 122, 128, 131; 238
 efflorescence, 113; 249
 electrical charges, 133; 135
 electrolytes, 58; 88, 96–102, 108–109
 electronegativity: table, 136
 electron patterns, 73–74
 electron-volt, 77; 179
 element(s), 38, 43, 44; 38, 44–46
 families of, 73–74; 164–172
 grouping of, 73, 79; 169
 periodic table of, 134–135
 endothermic solutions, 54; 86
 energy: in reactions, 88–89; 50–53
 randomness and, 33–34
 energy changes: in solutions, 53–54, 56; 83
 energy levels, 73; 148, 159, 163–167
 equations, 3, 95; 98–99, 201, 204,

214–218, 266–269, 296–297, 301–302
 equilibrium, 47–48; 27–28, 65, 10
 equilibrium systems, 64; 84–8
 ethyl alcohol, 85
 exothermic solutions, 54; 86

filtering, 13–15; 279–280
 fire polishing, 18–19
 flame tests, 67–68, 131; 276
 flow: rate of, 71
 flow chart: use of, 26
 formula(s), 110–112; 105, 207–209, 294, 303–304
 ions, 120
 freezing, 41–42; 252–254, 262
 gas(es): diffusion of, 100–101; 226, 227
 molar volume, 106–108; 240, 242
 pressure-temperature-volume relations, 102–105; 299–300
 tests for, 24–25
 gas burner, 5–7
 gas-liquid systems, 48
 glass: working with, 4, 18–20
 graphing, 28; 21–22

Handbook of Chemistry and Physics
 using, 35–36; 32, 34
 hardness, 86; 259–260
 heat energy: randomness and, 33, 34
 state and, 41; 70, 251, 254
 heat of neutralization, 88, 89; 10, 107
 heat of solution, 88–89
 hydration: water of, 113; 80
 hydron paper: use of, 37
 hydrogen: test for, 24
 hydrogen sulfide, 82–83
 hydrolysis, 63–65; 114–115
 hydronium ions: 68; 98
 hydroxide ions, 63; 106–107

indicators, 61–62; 107, 113
 indirect evidence, 66
 iodide: test for, 123
 ion(s): common, 63–65; 115–117, 275
 electrical charges, 133
 formulas, 120
 identification by color, 67–70
 oxidation numbers, 133
 reactions with metals, 94
 rearrangement of, 95
 tests for, 67–70

QD 33 L46 1973 LAB- MAN-
LEDBETTER ELAINE W
KEYS TO CHEMISTRY

39152266 CURR



000001247931

Date D	
DATE DUE SLIP	RETURN
	AUG 01 RETURN
DUE EDUC DEC 1 '85	NOV 22 '90
DEC 02 RETURN	NOV 16 RETURN
DUE EDUC MAR 03 '86	RETURN JAN 17 '01
MAR 03 RETURN	
FE DUE EDUC AUG 19 '86	
AUG 21 RETURN	
DUE EDUC MAR 19 '88	
MAR 18 RETURN	
DUE EDUC APR 19 '88	
APR 18 RETURN	
DUE EDUC JUN 07 '88	
JUN 09 RETURN	
JUL 26 '90	

F255

RECOMMENDED FOR USE
IN ALBERTA SCHOOLS

- ionic reactions, 118-121; 220
ionic solids; 86; 256, 265-266
ionization energy, 77-78; 164-168
iron: tests for, 129-130
- laboratory techniques, 23-27
lead: tests for, 128
Le Chatelier's Principle: common
 ion and, 64; 115-117
liquid(s): change to solid, 41
 electrical conductivity, 57-58;
 88, 96-102, 108-109
 handling, 8-10, 12-13, 17
 recovering solids from, 16-17
 volume of, 23-24
liquid-vapor systems, 47
litmus paper, 37
- magnetic force, 45-46
manganese, tests for, 130
matter: classification of, 37-38;
 37-56
 properties of, 39-40; 38-40
 maximum randomness, 42; 70, 254
measurement: density and, 31-32
melting: bonds and, 86-87; 188,
 194-196
 of solids, 41-42; 248, 250-251
meniscus, 10
mercury; tests for, 128-129
metal(s): reaction rate, 90; 98-100
 reactions with ions, 94; 219
metallic ion: color test for, 67-70;
 276-278
minimum energy, 42; 254
mixtures, 38, 43-44; 48-50
models, 66; 126-150
molar volume, 106-108; 240-242
molecular solids, 86; 257-258,
 263-265
molecular weight, 110-112; 82,
 213, 241-243
- nickel: tests for, 130
nitrate: test for, 124
nitrite: test for, 122
- observation, 21-22; 3, 5
orbitals, 73, 79; 157-163
oxidation numbers, 133, 137
oxygen: tests for, 25
- pH, 61, 63, 65, 128; 110-115, 275
paper chromatography, 71-72
periodic tables, 134-137; 170-171,
 184, 208
phosphate: test for, 123
physical properties, 39, 86-87; 38,
 40
positive ions, 70; 96, 273
potassium: test for, 131
precipitates: filtering, 13-15
 as tests, 123, 125, 128-131; 274-
 275
predictions: testing, 109
pressure: gas volume and, 102-103;
 234-235, 299-300
procedures, 5-20, 126-127
- qualitative analysis, 126-131; 273-
 281
qualitative observations, 21; 3-5
- randomness, 33-34, 56; 83, 266
reaction(s): complete, 97
 energy in, 88-89
 ionic, 118-121; 220
 promoted by water, 52
 quantitative study of, 97-99
 types of, 94-96; 218-220
reaction rate(s), 98-100
 concentration and, 91-92
 predicting, 8
 temperature and, 92
reagent bottles, 10-13
rearrangement reactions, 95; 219
recording observations, 2
residues, 14-15
rod: glass, 4, 18-19
- S.T.P., 106; 230
safety, 2, 4
salt(s): 59-60; 108-110, 114-115
scientists: work of, 1; 127
silver: test for, 128
sodium: test for, 131
solid(s): change to liquid, 42; 248,
 250-251
 measuring volume of, 31-32
 melting and freezing of, 41-42;
 248, 250-251
 recovering from liquids, 16-17
 taking from bottles, 10-11
solid-gas systems, 48
solid-liquid systems, 48
solubility: bond types and, 87; 263-
 266
solution(s); energy changes in, 53,
 56; 83
 ionic reactions in, 118-121
 process of, 21-22, 33-34, 84-85;
 62-63, 79-81, 265
 saturation, 55-56; 84-85
 water, 118-121; 77
 solvents, 72, 84-85; 76, 266
 standard temperature and pressure,
 106; 230
 states of matter, 37; 40-44
 changes of, 41-42
 heat energy and, 41; 70, 251, 254
 strontium: tests for, 130-131
 sulfate: test for, 123
 sulfide: test for, 122
 sulfite: test for, 122
 synthesis reactions, 94-95; 218
 systems: open and closed, 47-48
- technique, 2, 5-20, 23-27
temperature: gas volume and, 104-
 105, 235-236, 299-300
 reaction rate and, 91-93; 98-100
 temperature-time relations, 29-30
theories, 1; 150
tubing, 4, 18-20
- unknown(s): cations in, 81, 83
 determining formula of, 110-112
 identification of, 126-131
- van der Waals forces, 86; 194-196,
 257
vapor pressure, 50, 113, 132; 232,
 299
variables, 27-30; 20-24
volume: estimating, 12-13, 23-24
 gas pressure and, 102-103; 226,
 234-235, 299-300
 gas temperature and, 104-105;
 226, 235-236
 measuring, 8-10
 of liquids, 23-24; 63
 molar, 106-108; 240-242
- washing residues, 14
water: chemical properties of, 51-
 52; 73-74
 cooling, 29-30
 in crystals, 113-114; 80
 ionic reactions in, 118-121
 mass-volume relationship, 24
 promotes reaction, 52; 73-74
 reactions with, 51-52; 73-74
 solution process in, 84-85; 62-
 63, 79-81, 265
 vapor pressure, 132; 232, 298
 weight: measuring, 7-8
- zinc: test for, 130

B 21572

ADDISON-WESLEY PUBLISHING COMPANY

